Thesis submitted for Degree of Ph.D.

Degree of Ph.D. conferred

28th June, 1928.

AMPHOLYTES IN SOLUTION.

THE SOLUBILITIES OF SOME SALTS AS A MEASURE OF HYBRID ION FORMATION IN A SERIES OF AMINO ACIDS.

Thesis
presented for the Degree of Doctor of Philosophy

by

Walter Stoddart Macfarlane, B.Sc.

Chemistry Department,
University of Edinburgh.

May. 1928.

TABLE OF CONTENTS.

GRAPHS.

The classical theory of amphoteric electrolytes was developed in 1904 by Walker (Proc. Roy.Soc., 73,155). The dissociation of an ampholyte H.X.OH will give the ions H^+, OH^-, XOH^- and HX^+. If the concentrations of these ions be a,b,c and d Walker showed that their concentrations are given by the equations

$$(1)\quad a = \sqrt{\frac{K + k_a u}{1 + \frac{k_b}{K} u}}$$

$$(2)\quad b = \frac{K}{a}$$

$$(3)\quad c = k_a u \quad a$$

$$(4)\quad d = \frac{k_b}{K} u \quad a$$

where k_a = acidic dissociation constant

k_b = basic dissociation constant

K = dissociation constant of water

u = total concentration of unionised forms of the ampholyte.

Walker gave the mathematical treatment for finding the concentrations of these ions in any solutions. In this treatment it is unnecessary to consider whether the undissociated molecules exist as the hydrated form (e.g. in the case of glycine) $NH_3OH.CH_2.COOH$; the anhydrous form $NH_2.CH_2.COOH$, or in the form of the inner salt $\overline{NH_3.CH_2.COO}$. The form of inner salt composed of two molecules is ruled out on account of molecular weight determinations/

determinations which show that glycine exists chiefly as single molecules.

The further possibility of the simultaneous splitting off of a hydrogen ion and a hydroxyl ion leaving the molecule $^{+}NH_3.CH_2.COO^{-}$ was recognised by Bredig in 1894 (Z.physikal.Chem.13,323) who in a footnote mentions that betaine "is an inner salt and must therefore have on the same molecule a positive and a negative charge which neutralise each other. It would be interesting" he adds "to find out over what length of chain this neutralisation continues. An optical orientation of the betaine solution in an electric field could not be detected." In 1897 Küster (Z.anorg.Chem.13,135) in a paper dealing with the use of methyl orange advanced the theory (likewise in a footnote) that the red form of this indicator is the molecule $^{+}NH(CH_3)_2.C_6H_4.N_2.C_6H_4SO_3^{-}$. This species he referred to as a "Zwitterion" which name has been largely used in German literature since the recognition in recent years of the importance of this type of molecule. The word is literally translated into English as "hermaphrodite ion" but it may appropriately be referred to as "hybrid ion". Walker (Proc.Roy.Soc.78A,143) recognised the possibility of the existence of such ions but considered it was a complication unnecessary to explain the facts then known concerning ampholytes.

In/

In 1923 Bjerrum (Z.physikal.Chem.,104,147) advanced the theory that the "undissociated" part of an ampholyte in solution (i.e. the part that does not occur as cation or anion) exists practically completely in the form of hybrid ions. Fortunately this assumption does not in any way conflict with the older methods of calculating the equilibrium between ,anion, cation and undissociated molecules so that the work of Winkelblech,Walker,Bredig,Michaelis and others loses none of its importance. According to Bjerrum's theory the dissociation of the carboxyl group of an amino acid occurs thus

$$\underset{\text{concentration} = A^{+}}{{}^{+}NH_3.R.COO.H} = \underset{\text{concentration} = A^{+-}}{{}^{+}NH_3.R.COO^{-}} + H^{+}$$

and the corresponding acid dissociation constant is

$$K_s = \frac{A^{+-} \cdot H^{+}}{A^{+}}$$

The dissociation of the amino group is

$$\underset{\text{concentration} = A^{-}}{H_2O + NH_3.R.COO^{-}} = {}^{+}NH_3.R.COO^{-} + OH^{-}$$

and the corresponding basic dissociation constant is

$$K_B = \frac{A^{+-} \cdot OH^{-}}{A^{-}}$$

According to the older treatment the acid and basic dissociation constants are given by

$$k_a = \frac{A^{-} \cdot H^{+}}{A}$$

$$k_b = \frac{A^{+} \cdot OH^{-}}{A}$$

The relationship between the old and new dissociation constants/

constants is thus apparent. For, substituting A for A^{+-} we have

$$K_S = \frac{A^{+-}.H^+}{A^+} = \frac{A.H^+}{A^+} = \frac{A^+.OH^-.H^+}{A^+.k_b} = \frac{K_{H_2O}}{k_b}$$

and

$$K_B = \frac{A^{+-}.OH^-}{A^-} = \frac{A.OH^-}{A^-} = \frac{OH^-.A^-.H^+}{A^-\ k_a} = \frac{K_{H_2O}}{k_a}$$

Bjerrum gives a large number of reasons why the values of the acidic and basic constants arrived at on this theory are more plausible than the older values. The older values show that the amino-carboxylic acids have acid dissociation constants of from 10^{-8} to 10^{-10} whereas the carboxylic acids have constants from 10^{-2} to 10^{-5}. Mere substitution of NH_2 for H is unlikely to produce any such huge decrease. On the new theory the acid constants K_S have values $10^{-1.5}$ to $10^{-3.5}$ - somewhat higher than the carboxylic acids, asis to be expected from the fact that the dissociating molecule is $^+H_3N.R.COO.H$ which, having a +ve charge, will be more readily able to dissociate a H^+ ion than the uncharged HR.COO H. For the aminosulphonic acid taurine $NH_2.C_2H_4.SO_3H$ the value of k_a is $10^{-8.8}$ which is not in agreement with the fact that all sulphonic acids are strong acids. The new theory gives $K_S = 1$ as is to be expected for a strong acid. Similar evidence/

evidence is obtained from otheracids and also from a consideration of the basic dissociation constants. Further support for this theory is got from the "formoltitration" for amino acids, in which formaldehyde is added and the titration can then be done with phenolphthalein as indicator. The addition of formaldehyde causes the buffer action in the region of $pH = 10$ to disappear enabling a sharp end point to be obtained. According to the old theory the buffer action in this region is due to the COO H group but according to the new theory it is due to the NH_2 group. The new theory is thus in agreement with this experiment as it is well known that the formaldehyde acts on the amino group not on the acid group.

The estimation of the ratio $\frac{A^{+-}}{A}$ i.e. the ratio of hybrid ions to amino acid molecules is one of great difficulty.

Two methods were suggested by Bjerrum.

Firstly, in the case of an ampholyte which undergoes a colour change in one of its modes of dissociation, the hybrid ion will have the colour of the anion or cation while the undissociated molecule will have the other colour. By measuring the extinction coefficients for a certain wave length in strongly acid, and strongly alkaline solution the values for cation and anion may be found. An estimate of the ratio of hybrid ion to amino acid

can/

can then be made from a measurement of the extinction coefficient of the pure ampholyte in solution. This met hod does not appear to have been used.

A second method of estimating $\frac{A^{+-}}{A}$ depends on the solubility of the ampholyte in neutral salt solutions. The effect of neutral salts on the solubility of $NH_2R\ COO\ H$ should not be any greater than in the case pf other non-electrolytes. The effect of neutral salts on the solubility of $^{+}NH_3,R.COO^{-}$ in consequence of the electric charges should be to lower the activity and increase the solubility as is always the case when a salt is dissolved in another salt without a common ion. If the length of the carbon chain between the charges were very great the solubility effect would approximate to that of a common uni-univalent salt. The effect must be much less however in ordinary ampholytes owing to the neutralising action of the charges on each other. Bjerrum quotes one case in which this was found to be so. The solubility of methyl orange in potassium chloride solutions of different concentrations gives a value of k, the activity constant, about 1/3 of that which is found in the case of uni-valent salts.

The converse should also hold. The solubility of a salt in an aqueous solution of ampholytes should be greater than in water owing to the electric charges onthe ampholyte lowering the activity of the saturating salt and hence increasing its solubility.

Again the effect cannot be so great as in the

case/

case when the solvent solution is a uni-univalent salt of equivalent concentration. One figure is given by Bjerrum in support of this. The solubility increment of croceo-cobaltinitrate in .10 N Glycine solution is about 1/3 of the increment in the case of uni-univalent salt solutions.

These two figures are the only physico-chemical evidence for the existence of hybrid ions in the solution of an ampholyte. The arguments from analogy by which Bjerrum's theory seems so convincing lead to the conclusion that there are a preponderance of hybrid ions in such solutions. If so, then a certain effect on solubilities must follow, the magnitude of which however it is not easy to predict. It therefore seemed to the author desirable to carry out a more extensive investigation of this effect.

The objects of such a research would be to ascertain

(1) whether the effect is common to all salts

(2) Whether it is common to all ampholytes.

(3) What the magnitude of the effect is in different cases.

and (4) Whether the magnitude of the effect varies with increasing separation of the charges on the ampholyte or in any other regular manner.

There/

There remains the question of which is the best method to adopt:- the determination of the solubility of the ampholytes in salt solutions, or the determination of the solubility of salts in ampholyte solutions. As it was desired to investigate the effects with increasing length of chain, and as the only series of ampholytes considered for this purpose was the series of aliphatic amino acids it was necessary to use the method of solubility of salt in ampholytes. The great solubility of these ampholytes in water renders them unsuitable for solubility measurements where general, not specific effects are looked for. It was then decided to prepare the first four members of the ω-amino acids of the fatty series. i.e. α amino acetic acid (glycine) β amino propionic acid (β alanine), γ amino butyric acid and δ amino valeric acid. In addition it was decided to make the investigation more complete by including an ampholyte with a strongly acidic group, -β amino ethane sulphonic acid (taurine) and one with a strongly acidic and strongly basic group β-trimethyl-amino-ethane sulphonic acid (taurobetaine).

The salts used for the solubility measurements should fulfil the following requirements. They should be very sparingly soluble; (in order that the general and not specific solubility effects should be/

be found), and they should yet be sufficiently soluble to enable accurate determinations to be made in small quantities of solution. (This was necessary as, in some cases γ and δ amino acids, only a very small quantity was available. For this reason it was desirable to use a salt that could be estimated volumetrically, as the accuracy in small quantities of dilute solution is then greater than by gravimetric methods. For these reasons two salts were selected, one uni-univalent salt, silver bromate, and one uni-divalent salt, lead bromide, both of which can be rapidly and accurately estimated using I c.c. of the saturated solution. In the course of the research it was found necessary to introduce other salts for reasons which will appear. The salts chosen have the same requirements. They were thallous bromate, thallous thiocyanate and calcium iodate.

Glycine.

The glycine was prepared by the common method. Methylene amino acetonitrile was prepared from formaldehyde, ammonium chloride and sodium cyanide. It was twice recrystallised from water. The hydrolysis of methylene amino acetonitrile was effected by hydrobromic acid which was used in preference to hydrochloric acid owing to the greater solubility of ammonium bromide in methyl alcohol. 100 gms of methyleneamino-acetonitrile and 1 litre of 40% hydrobromic acid were heated for 3 hours under a reflux. Formaldehyde and dilute acid were then distilled off until separation of ammonium bromide caused bumping. The contents of the flask were then filtered hot and washed with a little water. The filtrate was then evaporated to almost complete dryness leaving a residue of glycine hydrobromide and ammonium bromide. This was dissolved in 1 litre of cold methyl alcohol and 100 c.c. of pyridine were added with vigorous shaking. Free glycine was precipitated slowly. After standing overnight the precipitate was collected on a filter and washed repeatedly with methyl alcohol. It was then dissolved in hot water and boiled with animal charcoal, filtered and a large excess of methyl alcohol added. Glycine was precipitated. It was recrystallised three times from aqueous alcohol. The m.p. was 226° approx. but it decomposes in melting. The product was free from halogen and was not further analysed.

β amino propionic acid.

β-amino propionic acid has been prepared by

(1) Action of ammonia on β iodopropionic acid (Mulder Ber.,9,1903).

(2) reduction of β oximino propionic acid with sodium amalgam (Pechmann; Annalen 264 288).

(3) reduction of cyanacetic acid with zinc and Sulphuric acid (Enzel; Ber.,8,1597).

(4) action of alcoholic ammonia on acrylic ester (Werder; Gazzetta,19 438).

(5) action of alkaline hypobromite on succinimide.

The last method is the most useful for preparing the acid. It was used by Hoogewerff and van Dorp (Rec., Trav. Chim. 10 5) and modified by Lengfield and Stieglitz (Amer. Chem. J.,15 508) and by Hale and Honan (J. Amer. Chem. Soc.,41 774). The purities of the acids prepared by these authors seem to be of different degrees. Thus Mulder (Ber.,9 1903) gives m.p. 180°; Hoogewerf and van Dorp m.p. 196° and Lengfield and Stieglitz m.p. 206°. The method here used is similar to that of Lengfield and Stieglitz as far as the stage of the ester hydrochloride. There after the method used for obtaining the free acid in the pure state is different.

30 gms. of Bromine were dissolved in 800 c.c. of a 10% Solution of potassium hydroxide. 20 gms. of succinimide were added and the solution slowly heated to 60° and kept between 60° and 70° for 2 – 3 hours. The/

The solution was then cooled in ice, acidified with hydrochloric acid, and evaporated to dryness on the steam bath. The residue which consisted chiefly of potassium chloride along with the hydrochloride of β amino-propionic acid and some succinic acid, was dried completely in a vacuum desiccator. It was then placed in a continuous extraction apparatus of Paterson's design and exhausted with absolute alcohol. This dissolved out theorganic matter leaving the potassium chloride which is only very slightly soluble in absolute alcohol.

For the isolation of the pure amino acid various processes involving the use of lead, barium and silver compounds have been tried. Of these the simplest is that described by Lengfield and Stieglitz of esterifying the amino acid hydrochloride, hydrolysing the ester with baryta, precipitating the baryta with excess of sulphuric acid, freeing the solution from excess of sulphuric acid by boiling with lead oxide, and from the final solution crystallising the free amino acid. This method has the obvious disadvantage that the final product will not be free from chlorine, owing to the appreciable solubility of lead chloride. Several attempts were made to prepare the acid by this method but in only one case could the product be induced to crystallise and in this case it was found to be impossible to free the crystalline/

crystalline hygroscopic product from chloride. After a few trials the following method was found to give a pure white non-hygroscopic crystalline product.

The extract containing the amino-acid hydrochloride (already partially esterified by the hot alcohol) along with some succinic acid and some potassium chloride was placed in a flask and dry hydrochloric acid passed in until the alcoholic solution was saturated. When the alcohol boiled, owing to the great heat of solution of HCl in alcohol, the flask was cooled in ice and more HCl passed in until the solution would absorb no more HCl in the cold. The liquid was allowed to stand overnight and decanted from KCl which was completely precipitated. The alcohol was evaporated off. The residue consisting of the hydrochloride of β-amino-propionic ester and succinic ester was dissolved in hot water and boiled with excess of barium hydroxide for 1 hour, to hydrolyse the esters. The solution was cooled and filtered; a slight excess of hydrochloric acid was added, and the solution evaporated to dryness. The residue contains barium chloride, succinic acid and the hydrochloride of β-amino propionic acid. It was washed repeatedly with anhydrous ether to remove the succinic acid, and then repeatedly with absolute alcohol to dissolve the hydrochloride. The alcoholic extracts were mixed with water and evaporated to dryness. The water waa added to prevent the formation of ester/

of ester which takes place when the hydrochloride is heated with absolute alcohol. By extracting in the cold very little, if any, esterification takes place. The final residue of the hydrochloride of β-amino propionic acid was white and crystalline and its solution in distilled water gives no precipitate with sulphuric acid (absence of barium) or with barium nitrate (absence of sulphate).

The hydrochloride was dissolved in water and excess of freshly precipitated silver carbonate was added. Into the filtered solution was passed hydrogen sulphide to eliminate the silver and the filtrate from this was evaporated to dryness giving a crystalline residue which was recrystallised twice from aqueous alcohol. The acid is very soluble in water but apparently quite insoluble in all organic solvents. Yield 4 gms. from 20 gms. of succinimide. It melted with decomposition at 204°C (corr.).

Melting points given in the literature are

180° Mulder (Ber., 9 1903)

196° Hoogewerf and van Dorp (Rec.trav.Chim. 10 5)

206° Lengfield and Stieglitz (Amer.Chem.J. 15 508)

The analyses of the compound for nitrogen was done by the Micro Kjeldahl Method of Pregl using N/70 solutions of HCl and NaOH, and methyl red as indicator. 1 c.c. of the acid solution corresponds to .0002 gm of nitrogen.

Results/

Results.

1. Wt. of acid .0313 gm.
Vol. of acid used up = 24.40 c.c.

$$\therefore \% \text{ of } N = \frac{24.40 \times .0002}{.0313} \times 100 = 15.60\%$$

2. Wt. of acid .0285 gm.
Vol. of acid used up = 22.30 c.c.

$$\% \text{ of } N = \frac{22.3 \times .0002}{.0285} \times 100 = 15.65\%$$

% of N (found) = 15.60%
15.65%

$$\% \text{ of } N \text{ (calc.)} = \frac{14}{89} = 15.73\%.$$

γ- amino butyric acid.

γ amino butyric acid has been obtained by the following methods.

1. by the oxidation of piperidino-formic ester
2. from γ-chlor butyronitrile by the action of potassium phthalimide and subsequent hydrolysis.
3. from ethylene dibromide by a combination of the phthalimide and malonic ester syntheses.
4. by the hydrolysis of pyrollidone.

Methods 1, 3 and 4 have been tried in the attempt to obtain a good yield.

γ amino butyric acid from piperidine.

This method was described by Schotten who obtained an acid which he named piperidinic acid. (Ber. 16 646). It was also used by Gabriel, who identified the acid obtained with γ-amino butyric acid synthesised by him by method (2). (Ber. 23 1771.)

Preparation of Piperidino-formic ester.

1 mol. of ethyl chloroformate dissolved in ether was slowly dropped into 2 mols. of piperidine in a large excess of ether kept cool by a freezing mixture. A violent reaction took place. Piperidine hydrochloride separated as a bulky solid which was filtered off and the solution evaporated. Piperidino-formic ester was left as a colourless oil which distilled unchanged at 211° C.

$$2\ C_5H_{10}NH + Cl.COO\ Et = C_5H_{10}NH.HCl + C_5H_{10}.H.COO\ Et$$

The yield was 90% of the theory.

Oxidation of piperidine formic ester.

25 c.c. of the ester were slowly dropped into 50 c.c. of cooled fuming nitric acid. The solution was poured into water and extracted with ether; on evaporation ~~with~~ of the ether a viscous pale yellow liquid remained (about 3 c.c. from 25 c.c. of the ester). This was put in a Carius tube with 15 c.c. of concentrated hydrochloric acid and heated to 140° for 2 hours. The tube was opened and the contents evaporated on the steam bath. The residue was dissolved in water, boiled with silver carbonate, filtered, treated with hydrogen sulphide, filtered, and evaporated to small volume. On adding a large excess of methyl alcohol a small quantity of a syrup was precipitated. After prolonged washing with ether, the syrup was dissolved in a very little water and excess of methyl alcohol added. A very small quantity of a light brown crystalline substance appeared, which melted about 180°. (Gabriel 184°. M.P. of the pure acid obtained by other methods 204°- 206°). It was recrystallised several times but in all cases it was extremely difficult to induce crystallisation and the product was always slightly coloured, although boiled with animal charcoal repeatedly. The M.P. was raised to 190° but by this time the yield was too small to be of use. An analysis however showed approximately/

approximately correct nitrogen content,

wt. of acid = .0285 gm.
vol. of N/70 acid = 19.20 c.c.

(found) $\therefore$ $\%N = \frac{19.20}{.0285} \times .0002 \times 100 = 13.48\%$

(calc.) $\%N = \frac{14}{103} \times 100 = \underline{13.60\%}$

γ amino butyric acid from ethylene dibromide.

Preparation of β-bromethyl phthalimide.

30 gms. of potassium phthalimide (obtained by the mixing of alcoholic caustic potash with an equi-molecular amount of phthalimide in boiling alcoholic solution), and 100 gms. of ethylene dibromide were placed in a flask fitted with a reflux condenser and with an efficient stirrer, the stem of which was water cooled above the point at which it entered the flask. The mixture was heated in an oil bath kept at 180-190° for 24 hours with vigorous mechanical stirring. The excess of ethylene dibromide was distilled off and the bromo-ethylphthalimide was extracted from potassium bromide by refluxing with 100 c.c. of absolute alcohol for 1½ hours. After distilling off the alcohol the residue was refluxed with 100 c.c. of carbon disulphide which left undissolved the diphthalimidoethane. After distilling off the carbon disulphide there remained about 15 gms. of light brown crystals of β-bromoethyl phthalimide which melted at 78°C. A recrystallised portion was colourless and melted at 82°C.

Condensation/

Condensation of β-bromoethylphthalimide with sodio-malonic ester.

1.2 gms. of sodium were dissolved in 12 c.c. of absolute alcohol and 10 gms. of malonic ester added. The clear solution was mixed with 10 gms. of β bromo-ethyl phthalimide and refluxed for 4 hours, sodium bromide separated in the flask. The alcohol and unchanged malonic ester were distilled off in steam and the residual oil extracted with ether. After separation of the ether, the syrupy residue (about 5 gms.) was heated in a sealed tube with 15 c.c. of concentrated hydrochloric acid to 180° for 3 hours. On opening the tube the contents were filtered from phthalic acid and the residue consisting of the hydrochloride of γ-amino butyric acid was dissolved in water, boiled with silver carbonate, filtered, treated with H_2S, filtered, evaporated to dryness, washed with ether, recrystallised from aqueous alcohol. Yield about ½ gm. The product was pure white and crystalline and melted at 202° C (corr.).

γ amino butyric acid from pyrrolidone.

This method is apparently the best method of preparing γ-amino butyric acid; as it gives almost quantitative yields. Pyrrolidone is obtained in good yield according to Tafel and Stern (Ber., 33 2224) by the electrolytic reduction of succinimide in sulphuric acid solution in a lead cathode vessel. This reduction according to Tafel and Stern is prevented by traces/

traces of other metals. Whether the vessel used was impure or in some other way the exact conditions were not fulfilled, I obtained only small yields of pyrrolidone, although in all cases some was undoubtedly formed as evidenced by the characteristic smell. 34 gms. of succinimide were dissolved in 50% sulphuric acid to make 113 c.c. of solution in a beaker shaped lead cathode of 60 mm. internal diameter and 80 mm. height. A porous pot of 31 mm. outer diameter was used to contain the anode solution. The anode consisted of a lead rod kept cool by a spiral coil through which cold water from the tap was circulated. A current of 14 amperes, was passed for 7 hours, (current density 120 amp., ratio of cathode surface to cathode volume 1:1). The liquid in the cathode compartment was then diluted with water, neutralised with barium carbonate and filtered and the water distilled off under reduced pressure. A yellow oil remained which distilled at 250°. This was pyrrolidone. The pyrrolidone was boiled with concentrated hydrochloric acid for 5 days and the product evaporated to dryness. The partially crystalline residue was dissolved in water, treated with excess of silver carbonate and filtered; the filtrate was treated with H_2S and filtered; and the filtrate was evaporated to dryness. The residue was dissolved in the minimum quantity of water, about 4 times the bulk of ethyl alcohol/

alcohol was added and ether until a ppt. just appeared. On standing a white crystalline solid separated which was filtered off and recrystallised twice from pure water. About 1½ gm. was obtained. The M.P. was 202.5° (corr.).

The Analyses were done by the Micro-Kjehldahl method

Analyses.

I. wt. of acid .0322 gm.
Vol. of N/70 acid used up = 21.82 c.c.

$$\therefore \% \text{ of } N = \frac{21.82}{.0322} \times .0002 \times 100 = 13.55\%$$

2. wt. of acid .0350
Vol. of N/70 acid used up 23.80 c.c.

$$\% \text{ of } N = \frac{23.80}{.0350} \times .0002 \times 100 = 13.60\%$$

$$\% \text{ of } N \text{ (calc.)} = \frac{14}{103} \times 100 = \underline{13.60\%}$$

δ amino valeric acid.

δ amino valeric acid was obtained by Schotten (Ber.,17,2546 and 21,2240) who oxidised the N-benzoyl derivative of piperidine with potassium permanganate in presence of dilute sulphuric acid. The product was heated in sealed tubes with concentrated hydrochloric acid and from the resulting solution he worked up an acid which he named "homopiperidinic acid". Gabriel (Ber.,23,1769) used a malonic ester synthesis. From γ brom-propyl phthalimide and sodio-malonic ester he synthesised γ amino butyric acid and showed that it was identical with Schotten's acid. This was the method I adopted for its preparation.

Preparation of γ Brom Propyl phthalimide.

This has been prepared by Gabriel (Ber.,21 2671). His method was modified in several respects and the method here used was identical with that used in the preparation of β-brom ethyl phthalimide (see p. 18). 50 gms. trimethylenbromide and 20 gms. potassium phthalimide were placed in the flask; the stirrer was started and the whole heated to 180° for 12 hours. The excess of trimethylene bromide was then distilled off/

off and the residue was extracted with boiling carbon disulphide which left about 10 gms. of diphthalimido-propane. The carbon disulphide was distilled off and the syrupy residue then set to a hard mass which was recrystallised twice from aqueous alcohol. The melting point of the product was 70° C. Gabriel gives 72° as the M.P. of pure γ Brom-Propyl phthalimide. The yield was about 4 gms. The process was repeated three times but in no case was a larger yield secured, a large amount of diphthalimido-propane being obtained in all cases.

Condensation of α brom-propyl phthalimide with sodio-malonic ester.

1.2 gms. of sodium were dissolved in 12 c.c. of absolute alcohol and 10 gms of malonic ester added. To the clear solution was added 10 gms γ brom-propyl phthalimide and the mixture refluxed for 4 hours. The excess of malonic ester and alcohol were distilled off in steam and the oil which remained at the foot of the flask was extracted with ether. After evaporation of the ether the syrupy residue set to a crystalline mass (about 4 gms.) which was placed in 2 portions in sealed tubes with 15 c.c. of concentrated hydrochloric acid in each. The tubes were heated to 180° for 3 hours. On opening the tubes the contents were found to be almost solid, owing to separated phthalic acid. This was filtered off/

off and the residue evaporated to dryness; the residue consisting chiefly of the hydrochloride of amino butyric acid was repeatedly washed with ether to remove any phthalic acid, dissolved in water, boiled with animal charcoal, filtered, boiled with silver carbonate, filtered, boiled with H_2S filtered and evaporated to dryness; it was thenwell washed with ether and recrystallised from aqueous alcohol twice. Yield about ½ gm.

The M.P. was 165 °*.

Analysis by Micro Kjehldahl method

Wt. of acid .0275 gm.

Vol. of N/70 acid = 16.40 c.c.

$$\% N = \frac{16.40}{.0275} \times .0002 \times 100 = 11.93\%$$

$$\%N \text{ (calc.)} = \frac{14}{11.7} = 11.97\%$$

*This is 10° higher than that found by Gabriel and 7° higher than that found by Schotten.

Taurine.

β amino-ethane sulphonic acid.

This acid was made by the method recently described by Marvel, Bailey and Sparberg (J.Amer.Chem. Soc., 1924,49,1833). Their method consists in allowing sodium β -bromo-ethane sulphonate to react slowly in the cold with ammonia in concentrated aqueous solution, evaporating the residue and re-crystallising from aqueous alcohol to separate taurine from sodium bromide.

Preparation of sodium bromo ethane sulphonate.

In a 3 litre flask fitted with reflux condenser, a mechanical stirrer and a tap funnel,were placed 300 gms. of ethylene dibromide, 600 c.c. absolute alcohol and 225 c.c. of water. The stirrer was started, the mixture heated to boiling and 60 gms. of anhydrous sodium sulphite dissolved in 225 c.c., of water was slowly added from the dropping funnel over a period of about two hours. The boiling was continued for 2 hours longer and the excess of ethylene bromide and alcohol were distilled off. The remaining solution was evaporated to dryness on the steam bath. The sodium β -bromo-ethane sulphonate was extracted from sodium bromide and sodium sulphite by boiling with 1 litre of 95% aqueous alcohol and filtering. On cooling a bulky precipitate of sodium 2-bromo/

2-bromo-ethane sulphonate was obtained, which contained a small quantity of sodium bromide but was not further purified. Yield 51 grams.

$$Br\ C_2H_4Br + Na\ SO_3Na = NaBr + Br\ C_2H_4SO\ Na.$$

Preparation of Taurine.

50 grams of sodium 2 bromo ethane sulphonate was dissolved in 1 litre of concentrated aqueous ammonia and allowed to stand for seven days in the cold. The solution was then evaporated to dryness; the residue was dissolved in a little hot water, boiled with animal charcoal, filtered and concentrated to about 50 c.c. About 200 c.c. of boiling alcohol were then added and the solution allowed to cool. On cooling a large quantity of crystalline precipitate separated, which contained a considerable amount of bromide. The process of dissolving in hot water, adding hot alcohol and cooling was repeated 3 times when a crystalline precipitate free from bromide was obtained. The product was recrystallised twice from pure water giving fine prismatic crystals.

$$Br\ CH_2.CH_2.SO_3Na + NH_3 = NH_2.CH_2.CH_2.SO_3H + NaBr$$

As usual the product was analysed by the Micro Kjehldahl method but it always gave low results. Even on very prolonged heating (3 days) with concentrated sulphuric acid and an equal weight of potassium bisulphate using mercuric sulphate, or potassium perchlorate/

perchlorate as catalyst, the substance was not completely decomposed. This was most readily seen when a copper salt was used as catalyst, as the solution then retained the blue colour of a complex copper salt even after boiling with concentrated sodium hydroxide.

Typical results were

%N 6.56; 10.71; 10.33; 7.61; 9.43;
10.48; 10.74; 8.21;

%N (calculated) = 11.20%.

The analyses for the sulphur content were done by a Micro-Carius Method, using 6 inch lengths of $\frac{1}{4}$" hard glass tubing sealed at one end. The weighed acid was introduced into such a tube, and 20 drops of fuming nitric acid added. The tube was then sealed and heated to 300° for 2 hours, cooled, and opened. The contents were rinsed out and evaporated twice with concentrated HCl to eliminate nitrate. The sulphate was then precipitated as $BaSO_4$ and weighed in Gooch Crucibles.

Results. (1) Wt. of taurine .0348
Wt. of $BaSO_4$ .0646

$$\% \text{ of S} = \frac{.0646 \times 32 \times 100}{.0348 \times 233.4}$$

$$= \underline{25.44}$$

(2) Wt. of taurine .0370
Wt. of $BaSO_4$ .0689

$$\% \text{ of S} = \frac{.0689 \times 32 \times 100}{.0370 \times 233.4}$$

$$= \underline{25.53}$$

% of S (calculated) = 25.60

An analysis for nitrogen was then made by the Dumas combustion method.

This was done in the ordinary type of combustion tube but a micro-nitrometer of 1 c.c. capacity was used forncollecting the nitrogen.

Result.

Wt. of taurine = .00700 gm.

Vol.of nitrogen = .670 c.c.

Temperature = 17°C

Pressure = 760 m.m.

$\therefore$ Vol. of N_2 at N.T.P.$= .670 \times \frac{273}{290}$

$\therefore$ % of N $= .670 \times \frac{273}{290} \times \frac{14}{11,200} \times \frac{100}{.007}$

$= 11.25\%$

% of N (calculated) $= \frac{14}{125} = 11.20\%$.

Taurobetaine.

Trimethylamino-ethane sulphonic acid.

This substance was apparently prepared in 1885 by James (J.pr.Chem. (2) 31, 418) who heated the trimethylamine salt of β chlorethane sulphonic acid with alcoholic trimethylamine solution in sealed tubes. He obtained a substance to which he ascribed the formula $(CH_3)_2N.C_2H_4.SO_2.OCH_3$ but which certainly from its properties is almost/identical with taurobetaine. The method is not a good one for the preparation of taurobetaine as the yields are small and the starting substance (β chlor-ethane sulphonic acid) is very difficult to prepare.

A better method for the preparation is found in the action of sulphur dioxide on neurine. (trimethyl-vinyl ammonium hydroxide) Schmidt and Wagner (Annalen 337 63) found that an aqueous solution of neurine absorbed sulphurous acid when its solution saturated with SO_2 stood in the cold for several days. On evaporation of the solution small white crystals which did not melt at 250°C were obtained. This method was used and was found to give a fair yield of very pure product.

Preparation of Bromocholine Bromide.

Equimolecular quantities of dry trimethylamine and ethylene dibromide were heated in sealed tubes to 70° for 12 hours. By this time the contents of/

of the tubes had completely crystallised. The crystals were washed twice with ether and recrystallised from absolute alcohol giving a fairly pure product. The yield was 90% of the theory. The bromocholine bromide was again recrystallised from alcohol and the M.P. was then found to be 231°. Renshaw (J.Amer.Chem. Soc., 34, 1618) found 235.5° and Bode (Annalen 267,268) found 230°.

Preparation of Neurine and Taurobetaine.

40 gms. pure bromocholine bromide were dissolved in water, cooled in ice, and shaken with 80 gm. of freshly precipitated silver hydroxide, until a filtered portion no longer reacted with silver nitrate. This method of preparing a solution of neurine is described by Meyer and Hopff (Ber., 54 2277). The solution which was then rapidly filtered contains the neurine, which is a very unstable substance. The aqueous solution was at once saturated with SO_2 and allowed to stand in the ice-house for 5 days when it was evaporated to dryness and the residue crystallised and recrystallised 3 times from aqueous alcohol. Yield 5 gms. from 80 gms. of bromocholine bromide.

$$Br.CH_2.CH_2.Br + N(CH_3)_3 = Br.CH_2.CH_2.N(CH_3)_3Br$$

bromocholine bromide.

$$Br.CH_2.CH_2.N(CH_3)_3Br + 2AgOH = CH_2:CH.N(CH_3)_3OH + 2AgBr + H_2O$$

neurine

$$CH_2:CH.N(CH_3)_3OH + H\,SO_3H = SO_3H.CH_2.CH_2.N(CH_3)_3OH$$

taurobetaine

The/

The substance does not melt at 350° and at higher temperatures it chars without melting.

Analyses. Nitrogen by Micro Kjehldahl Method.

(1) wt.of taurobetaine .0615 gm.
Vol.of N/70 acid 25.92 c.c.

$$\%N = \frac{25.92}{.0615} \times .0002 = 8.42\%$$

(2) wt. of taurobetaine .0626
Vol.of N/70 acid 26.30 c.c.

$$\%N = \frac{26.30}{.0627} \times .0002 = 8.40\%$$

% of N (calculated from formula $(CH_3)_3N.CH_2.CH_2.SO_3$)
= 8.38%

% of N (calculated from formula $(CH_3)_3OH.N.CH_2.CH_2.SO_3H$
= 7.54

Thus the formula is the anhydrous form.

Sulphur by Micro-Carius Method.

(1) wt.of taurobetaine .0358 gm.
wt.of $BaSO_4$ .0502 gm.

$$\%S = \frac{.0502 \times 32}{233.4 \times .0358} = 19.22\%$$

(2) wt.of taurobetaine .0301
wt.of $BaSO_4$ .0421

$$\%S = .0421 \times \frac{32}{233.4} \times \frac{100}{.0321} = 19.18\%.$$

% of S (calc.from formula $(CH_3)_3N.CH_2.CH_2.SO_3$)
= 19.16%

% of S (calc.from formula $(CH_3)_3.OH.N.CH_2.CH_2.SO_3H$)
= 17.30

Thus the formula is undoubtedly the anhydrous form.

EXPERIMENTAL I.

Preparation of the salts.

The saturating salts were all prepared by slow precipitation in dilute aqueous solution using the purest reagents obtainable.

Silver bromate was prepared by adding a dilute solution of potassium bromate (Merck's A.R.) to excess of a dilute solution of Silver Nitrate (A.R.). The precipitate was washed many times by decantation with distilled water and then filtered with suction and washed many times on the filter. The silver bromate was then recrystallised by passing steam into a suspension of it in water until the solution was boiling, then filtering hot. The silver bromate settled as a white minutely crystalline precipitate.

Thallous bromate was prepared in the same way from a recrystallised sample of thallous nitrate (B.D.H.) and Merck's potassium bromate. It was likewise a white minutely crystalline product.

Thallous thiocyanate was prepared from recrystallised thallous nitrate and ammonium thiocyanate (A.R.) It was obtained as thin plates with a greenish yellow lustre.

Lead Bromide was prepared from recrystallised lead acetate and potassium bromide (Kahlbaum's fur Analyse) It/

It was recrystallised by dissolving in hydrobromic acid (density 1.5) and slowly adding water to the solution. This precipitation was repeated twice. Lead bromide was precipitated as a very bulky but crystalline precipitate which after washing, filtration and drying was a heavy white powder which rapidly darkened on the surface when exposed to light.

Calcium Iodate was prepared from recrystallised calcium chloride and potassium iodate. As it was not considerably more soluble in hot than in cold water it was not recrystallised, but was very thoroughly washed with large quantities of hot water.

EXPERIMENTAL 2.

Preparation of the saturated solutions.

The solutions of the amino acids were made from accurately weighed quantities of the pure acids made up with conductivity water in standard flasks to the required volume. The saturation of the solution with the salt was carried out in 50 c.c., Jena glass conical flasks. About 20 c.c. of the solvent solution was placed in each flask; in those cases where little of the solvent was available i.e. γ-amino-butyric acid and δ-amino valeric acid, only 5 c.c. of the solution was used. An excess of the salt was then added. In some experiments the solution was then warmed for an hour to about 40°, before being stoppered/

stoppered, to ensure that supersaturation was reached from above as well as from below the saturation point. In other cases the flasks containing the cold solutions and salt were immediately stoppered. Rubber stoppers were used and after these had been tightly inserted the stopper and neck of the flask were dipped into molten paraffin wax just above its melting point. This on cooling formed a water-tight cover which did not crack on any occasion. The flasks were then placed in the shaking apparatus in the thermostat. This consisted of a horizontal shaft to which were attached four wooden discs about eight inches in diameter. Each of the discs could accomodate four 50 c.c flasks arranged radially; the necks of the flasks were attached to the discs by strips of brass which were held on to the discs by means of brass screw terminals. The shaft was rotated by means of a brass chain working on a pulley at one end of the shaft. The motive power was a hot air engine. This apparatus had the advantages of stirring the contents of the flasks very efficiently as they were turned end over end many times per minute, and also of very efficiently stirring the thermostat the regulation of which was easily maintained to .01 degree as read on a sensitive thermometer. The time allowed for saturation was always over 36 hours. It was found that there was no change in the solubility after 1 day of saturation and/

and after 7 days; so that 36 hours allows an ample margin of safety. It was also found that there was no consistent difference between the results obtained for the solutions where saturation was approached from above the saturation point and those where it was approached from under saturation. In the experimental results therefore no distinction has been made between these two methods. After at least 36 hours the flasks were removed from the shaker and placed on trays in the thermostat with their necks projecting above the water. After standing for half an hour the stoppers were dried and removed, care being taken that no wax fell into the solution. The excess of salt had settled out in heavy particles at the bottom of the flask and hence there was no need to filter the solutions. The required volume of the clear solution was simply pipetted off and analysed. Sometimes the first portion to be pipetted off gave incorrect results, perhaps owing to precipitation in the cold pipette but thereafter repeatable results were regularly obtained. No correction was applied for the difference in volume of the pipette at 15°C for which it was standardised and at 25°C the temperature at which it was used, the correction necessary being very much smaller than the experimental error of the methods used in analysis.

EXPERIMENTAL 3.

Estimation of the solubility.

The solubility of the bromate and iodate was estimated iodometrically. 1 or 2 c.c. of the saturated solution standing in the thermostat was pipetted off and added to an acid solution of potassium iodide and titrated with standard (about .01 N) sodium thiosulphate using starch as indicator. The method adopted was to dissolve about 10 gms of potassium iodide in about 50 c.c. of water in a 250 c.c. conical flask, add 10 c.c. dilute acid (hydrochloric or sulphuric 2 N) and 1 c.c. of starch solution. Usually a few drops of thiosulphate were required to decolourise the solution. 1 c.c (or 2 c.c.) of the saturated solution of the bromate or iodate was then added and the solution immediately titrated till colourless. The solution could at once be used again for further titrations. In this way several titrations could be done on each solution. In the case of the silver salts the end point was much better when sufficient excess of potassium iodide was used to dissolve the silver iodide. In the case of the thallous salt the end point was not easy to observe with accuracy owing to the separation of a yellow precipitate of thallous iodide not soluble in potassium iodide.

The centinormal thiosulphate was standardised each/

each day it was used against a standard solution of potassium dichromate and a standard solution of potassium bromate, as it was found slowly to decrease in concentration.

The solubility of thallous thiocyanate and lead bromide was found by adding excess of silver nitrate and back titrating with potassium thiocyanate. 1 c.c of saturated solution was pipetted off into a white porcelain basin, a measured volume, known to be excess, of standard (about .01 N) silver nitrate, and 2 c.c. of iron alum indicator were added and the solution titrated with standard potassium thiocyanate (about .01 N) with brisk stirring until a permanent brown colour appeared. Using bromides and thiocyanates there was no necessity to filter off the precipitated silver salts. The silver nitrate was standardised by means of weighed quantities of potassium bromide (Kahlbaum für Analyse).

In each case the first solubility estimations were done with the saturating salt in water alone with a view to testing the purity of the salts. This was done by two methods.

Firstly, the solubility of the salt was estimated using (a) a small excess and (b) a very large excess of the salt. Presence of a soluble impurity would lead to a marked difference between the two results. In no case was any consistent difference observed/

observed.

Secondly the solubility of the salt was measured in the usual way and then the excess of solution was poured off and more water added to the residue of salt in the flask, and the saturation repeated. The solubility was then found again, and in no case had it changed. This was taken as sufficient evidence of the purity of the salts.

In the tables are given the solubilities found for the several salts in the several amino acids. For comparison is given the behaviour of a typical non electrolyte (ethyl alcohol) and a typical uni-univalent salt (ammonium nitrate). In each case there is also given the ratio of the solubility in the given solution to the solubility in pure water $\frac{S}{S_0}$ and the ratio of the solubility increment in the given solution to the solubility increment in an equivalent solution of ammonium nitrate (r).

SOLUBILITY OF SILVER BROMATE.

The mean value of the solubility of silver bromate in water at 25° was found to be ·00827 gm. molecules per litre (= 1·951 gms/litre). This agrees with the figure obtained by Hill (J.Amer.Chem.Soc., 39,218) who using a very much more accurate method obtained 1·949 gms./litre at 25°. Noyes (Z.physikal. Chem.,6,246) at 24·5° obtained 1·911 gms/litre. The figure obtained by Longi (Gazzetta,13,87) namely 1·68 gms/litre appears to be too low.

Table 1..

Silver Bromate in Ethyl Alcohol.

Concentration of alcohol moles/litre. a	Concentration of salt moles/litre s	$\frac{S}{S_o}$
·000	·00827	1·000
·025	·00825	·997
·050	·00820	·991
·100	·00810	·980
·200	·00792	·958

Table 2.

Silver Bromate in Ammonium Nitrate.

a	s	$\frac{S}{S_o}$
·000	·00827	1·000
·025	·00900	1·088
·050	·00947	1·145
·100	·01008	1·219
·200	·01122	1·357

Table 3 .

Silver Bromate in Glycine.

a	s	$\frac{s}{s_o}$	r
.000	.00827	1.000	-
.025	.00854	1.033	.37
.050	.00870	1.052	.36
.100	.00909	1.099	.45
.200	.00972	1.175	.49

Mean = .42

Table 4 .

Silver Bromate in β-amino-propionic acid.

a	s	$\frac{s}{s_o}$	r
.000	.00827	1.000	-
.025	.00858	1.038	.43
.050	.00892	1.079	.54
.100	.00951	1.150	.68
.200	.01074	1.299	.84

Mean = .62

Table 5 .

Silver Bromate in γ-amino butyric acid.

a	s	$\frac{s}{s_o}$	r
.000	.00827	1.000	-
.025	.00880	1.064	.73
.050	.00950	1.150	1.02
.100	.01054	1.275	I.25
.200	.01160	1.403	I.13

Mean= 1.03

Table 6 .

Silver Bromate in δ-amino valeric acid.

a	s	$\frac{S}{S_o}$	r
.00	.00827	1.000	-
.10	.00950	1.149	.68
.20	.01067	1.290	.81

Mean = .75

Table 7 .

Silver Bromate in Taurine.

a	s	$\frac{S}{S_o}$	r
.000	.00827	1.000	-
.025	.00842	1.018	.20
.050	.00850	1.028	.19
.100	.00865	1.046	.21
.200	.00896	1.083	.23

Mean = .21

Table 8 .

Silver Bromate in Taurobetaine

a	s	$\frac{S}{S_o}$	r
.000	.00827	1.000	-
.025	.00842	1.018	.20
.050	.00853	1.031	.22
.100	.00871	1.053	.24
.200	.00905	1.094	.26

Mean = .23

Solubility of Lead Bromide.

The solubility of Lead Bromide in pure water at 25° was found to be ·0267 gm. molecules per litre (= 9·80 gms/litre), The value found by Lichty (J.Amer.Chem.Soc., 25 474) was 9·70 gms/litre.

Table 9.

Lead Bromide in Ethyl Alcohol

a	s	$\frac{S}{S_o}$
·000	·0267	1·000
·025	·0266	·996
·050	·0264	·989
·100	·0261	·978
·200	·0257	·963

Table 10.

Lead Bromide in Ammonium Nitrate.

a	s	$\frac{S}{S_o}$
·000	·0267	1·000
·025	·0287	1·073
·050	·0303	1·143
·100	·0335	1·258
·200	·0380	1·423

Table 11.

Solubility of Lead Bromide in the Amino Carboxylic Acids.

Solvent	s	$\frac{S}{S_o}$	r
Water	·0267	1·00	-
·200 N Glycine	·0475	1·78	1·84
·200 N β-Amino-propionic acid	·0730	2·73	4·10
·200 N γ Amino-butyric acid.	·0790	2·96	4·63

The solubilities shown in Table 11 are so obviously anomalous that the determination of the solubility at other concentrations of the amino acids was not proceeded with, nor was the solubility in δ-amino valeric acid found.

Table 12.

Solubility of Lead Bromide in Taurine

a	s	$\frac{S}{S_o}$	r
·000	·0267	1·000	-
·025	·0276	1·034	·45
·050	·0283	1·060	·44
·100	·0295	1·105	·41
·200	·0317	1·187	·44
		Mean =	·44

Table 13.

Solubility of Lead Bromide in Taurobetaine

a	s	$\frac{S}{S_o}$	r
·000	·0267	1·000	-
·025	·0276	1·034	·45
·050	·0285	1·067	·50
·100	·0300	1·124	·48
·200	·0327	1·224	·53
		Mean =	·49

Solubility of Thallous Thiocyanate.

The solubility of thallous thiocyanatein pure water was found to be .0147 gm. molecules per litre (3.86 gms./litre). This agrees fairly well with the value 3.90 gms./litre found by Noyes (Z.physikal. Chem., 6, 248).

Table 14.

Thallous Thiocyanate in Ethyl Alcohol.

a	s	$\frac{s}{s_o}$
.000	.0147	1.000
.025	.0147	1.000
.050	.0146	.993
.100	.0145	.986
.200	.0144	.980

Table 15.

Thallous Thiocyanate in Ammonium Nitrate.

a	s	$\frac{s}{s_o}$
.000	.0147	1.000
.025	.0158	1.075
.050	.0166	1.130
.100	.0180	1.225
.200	.0202	1.374

Table 16.

Thallous Thiocyanate in Glycine.

a	s	$\frac{s}{s_o}$	r
.000	.0147	1.000	-
.025	.0149	1.014	.18
.050	.0151	1.027	.21
.100	.0153	1.041	.18
.200	.0158	1.075	.20

Mean = .19

Table 17.

Thallous Thiocyanate in β-amino propionic acid

a	s	$\frac{s}{s_o}$	r
.000	.0147	1.000	-
.025	.0149	1.014	.18
.050	.0152	1.034	.26
.100	.0156	1.062	.27
.200	.0164	1.113	.31

Mean = .26

Table 18.

Thallous Thiocyanate in γ-amino butyric acid

a	s	$\frac{s}{s_o}$	r
.000	.0147	1.000	-
.025	.0151	1.027	.36
.050	.0154	1.048	.36
.100	.0162	1.102	.45
.200	.0174	1.184	.49

Mean = .41

Table 19.

Thallous Thiocyanate in δ-amino valeric acid.

a	s	$\frac{s}{s_o}$	r
.000	.0147	1.000	-
.100	.0161	1.095	.43
.200	.0170	1.156	.42

Mean = .42

Table 20.

Thallous Thiocyanate in Taurine.

a	s	$\frac{s}{s_o}$	r
.000	.0147	1.000	-
.025	.0149	1.014	.18
.050	.0152	1.034	.26
.100	.0157	1.068	.30
.200	.0166	1.130	.36

Mean = .27

Table 21.

Thallous Thiocyanate in Taurobetaine.

a	s	$\frac{s}{s_o}$	r
.000	.0147	1.000	-
.025	.0150	1.020	.27
.050	.0153	1.041	.32
.100	.0160	1.088	.39
.200	.0172	1.170	.45

Mean = .36

Solubility of Thallous Bromate.

The mean value of the solubility of thallous bromate in water at 25° was found to be .01238 gm. molecules per litre (= 4.11 gms. per litre).

Böttger (Z.physikal.Chem.,46,602) found the solubility at 19.9° to be .01043, while Noyes and Abbot (Z.physikal.Chem.16,130) found the solubility at 37.75° to be .02216. The value thus found at 25° is in agreement with these figures.

Table 22.

Thallous Bromate in Ethyl Alcohol

a	s	$\frac{S}{S_0}$
.000	.01238	1.000
.025	.01230	.994
.050	.01220	.986
.100	.01202	.971
.200	.01165	.941

Table 23.

Thallous Bromate in Ammonium Nitrate

a	s	$\frac{S}{S_0}$
.000	.01238	1.000
.025	.01304	1.053
.050	.01384	1.118
.100	.01546	1.249
.200	.01758	1.414

Table 24.

Thallous Bromate in Glycine.

a	s	$\frac{S}{S_0}$	r
.000	.01238	1.000	-
.025	.01255	1.014	.25
.050	.01266	1.023	.19
.100	.01312	1.059	.24
.200	.01380	1.114	.27

Mean = .24

Table 25.

Thallous Bromate in β-amino propionic acid.

a	s	$\frac{S}{S_0}$	r
.000	.01238	1.000	-
.025	.01255	1.014	.25
.050	.01268	1.024	.21
.100	.01325	1.070	.28
.200	.01404	1.134	.32

Mean = .26

Table 26.

Thallous Bromate in γ-amino butyric acid

a	s	$\frac{S}{S_0}$	r
.000	.01238	1.000	-
.025	.01258	1.016	.30
.050	.01273	1.028	.24
.100	.01332	1.076	.30
.200	.01423	1.150	.36

Mean = .30

Table 27.

Thallous Bromate in δ-amino-valeric acid.

a	s	$\frac{S}{S_V}$	
.00	.01238	1.000	-
.10	.01320	1.066	.27
.20	.01403	1.133	.32

Mean = .30

Table 28.

Thallous Bromate in Taurine.

a	s	$\frac{S}{S_o}$	r
.000	.01238	1.000	-
.025	.01247	1.007	.13
.050	.01255	1.014	.12
.100	.01300	1.051	.20
.200	.01356	1.096	.23

Mean = .17

Table 29.

Thallous Bromate in Taurobetaine.

a	s	$\frac{S}{S_o}$	r
.000	.01238	1.000	-
.025	.01250	1.009	.18
.050	.01271	1.027	.23
.100	.01312	1.059	.24
.200	.01378	1.113	.27

Mean = .23

Solubility of Calcium Iodate.

The solubility of calcium iodate in pure water at 25° was found to be .00790 gm.molecules/litre (= 3.08 gms./litre). The only values to be found in the literature are given by Mylius and Funk (Ber., 30, 1724) who found that when the solid phase was $Ca(IO_3)_2 \cdot 6 H_2O$ the solubility at 18° was 2.5 gms./litre and at 30° 4.2 gms./litre; when the solid phase was $Ca(IO_3)_2 \cdot H_2O$ the solubility was 3.7 gms./litre at 21° and 4.8 gms/litre at 35°.

Table 30.

Calcium Iodate in Ethyl Alcohol.

a	s	$\frac{s}{s_o}$
.000	.00790	1.000
.025	.00785	.994
.050	.00775	.981
.100	.00761	.964
.200	.00738	.935

Table 31.

Calcium Iodate in Ammonium Nitrate.

a	s	$\frac{s}{s_o}$
.000	.00790	1.000
.025	.00875	1.108
.050	.00945	1.196
.100	.01065	1.348
.200	.01249	1.580

Table 32.

Calcium Iodate in Glycine.

a	s	$\frac{s}{s_o}$	r
.000	.00790	1.000	-
.025	.00805	1.019	.18
.050	.00825	1.045	.22
.100	.00862	1.091	.26
.200	.00952	1.205	.35

Mean = .25

Table 33.

Calcium Iodate in β-amino-propionic acid.

a	s	$\frac{s}{s_o}$	r
.000	.00790	1.000	-
.025	.00811	1.027	.25
.050	.00830	1.051	.26
.100	.00866	1.096	.28
.200	.00950	1.203	.35

Mean = .28

Table 34.

Calcium Iodate in γ-amino butyric acid.

a	s	$\frac{s}{s_o}$	r
.000	.00790	1.000	-
.025	.00811	1.027	.25
.050	.00832	1.053	.26
.100	.00860	1.089	.25
.200	.00955	1.209	.36

Mean = .28

Table 35.

Calcium Iodate in δ-amino valeric acid.

a	s	$\frac{s}{s_o}$	r
.000	.00790	1.000	-
.100	.00883	1.118	.34
.200	.00964	1.220	.38

Mean = .36

Table 36.

Calcium Iodate in Taurine.

a	s	$\frac{s}{s_o}$	r
.000	.00790	1.000	-
.025	.00800	1.013	.12
.050	.00820	1.038	.19
.100	.00860	1.089	.25
.200	.00920	1.165	.28

Mean = .21

Table 37.

Calcium Iodate in Taurobetaine.

a	s	$\frac{s}{s_o}$	r
.000	.00790	1.000	-
.025	.00795	1.007	.06
.050	.00800	1.013	.06
.100	.00820	1.038	.11
.200	.00850	1.076	.13

Mean = .09

Discussion of Results.

From an inspection of Tables 1 to 37 , the first point observed is that in every case there is an increase in the solubility of thesalt in solutions of ampholytes. Six different ampholytes have been examined and five different salts have been used in conjunction with each and no exception has been found. Admittedly the results are not sufficiently numerous to enable a generalisation to be made with certainty but they strongly indicate that the general effect of ampholytes is to increase the solubility of salts.

When we come to examine the magnitude of this effect however we find great irregularities. In the fourth column of the tables, headed r , are given the ratios of the solubility increases found, to the solubility increase in equivalent solutions of ammonium nitrate. These figures give a rough idea of the effect. Consider first the case of silver bromate. These experiments were the first to be done. It became appaeent that the increase in the case of the β and γ amino-carboxylic acids was greater than had been expe cted. Thus, in the case of γ-amino butyric acid, the solubility increase was greater than in an equivalent solution of ammonium nitrate. This could not be due to the ionic effect which I was looking for, as it seems certain that a considerable portion of the field of each charge must be internal, thus reducing/

thus reducing the effects outside the molecule. Thus although the series of α, β and γ,-amino-carboxylic acids shows in a striking way the increase of the solubility effect with increase of length of chain ($r_\alpha = .42$ $r_\beta = .62$ $r_\gamma = 1.03$) it was thought that this could not really be the normal effect. This is borne out by the measurements with taurine and taurobetaine for which we have $r_\beta = .21$ and $.23$.

The other salt which it was originally intended to use, showed even more markedly a behaviour that cannot be attributed to an ionic effect. The values of r obtained were $r_\alpha = 1.84$ $r_\beta = 4.10$ $r_\gamma = 4.63$ obviously indicating some kind of specific interaction between the ampholyte and the salt. In this case again however the values with taurine and taurobetaine were more of the order to be expected, $r_\beta = .44$ and $.49$. It was noticed in the case of the solubility of the lead bromide that the crystalline appearance of the powder was lost and a bulky white substance appeared which covered the lead bromide and did not completely settle out of the solution. It was suspected that this was lead hydroxide; the solubility product of lead hydroxide can readily be exceeded in such solutions, but this is not to be expected in the cases where the acidic group is strong - taurine and taurobetaine. A similar effect may occur in the case of the silver salt with the exception that the silver hydroxide formed will readily form a complex with the amino group and will not be precipitated/

precipitated.

It was therefore decided to use as saturating salts the salts of soluble bases such as the calcium salts and the thallous salts. Calcium iodate, thallous thiocyanate and thallous bromate were selected as salts readily estimable in small quantities and having the requisite solubility.

An examination of these figures (Tables 14 to 37) shows much greater regularity. The value of r for thallous bromate is from ·17 to ·30. for thallous thiocyanate ·19 to ·42 and for calcium iodate ·09 to ·36. The case of calcium iodate in taurobetaine is apparently exceptional. The increase in solubility is still well marked but much smaller than in any of the other cases examined.

An explanation of this increase in solubility of salts in presence of ampholytes must now be sought. Solubility increase of salts is commonly due to the presence of ions in the solvent medium lowering the activity of the ions of the saturating salt. Now an ampholyte solution certainly contains ions according to the old and new theories alike. Only the real ions (the ions which move) are taken into account in the older theory. Is the presence of these ions sufficient to account for the observed increases? An approximate calculation shows that this is not so. In table 38 are given the values of the molecular conductivities of/

of glycine and β-amino propionic acid.

Table 38.

Molecular Conductivities of Ampholytes

Dilution	Molecular Conductivity of	
	Glycine*	β-amino-propionic acid**
4	·227	–
8	·236	·13
16	·246	·12
32	·257	·14
64	·275	·17
128	·306	·30
256	–	·25

* Ostwald (J.pr.Chem., 32, 369)
** Bork (Z.physikal.Chem., 129, 60)

Now, the molecular conductivities at infinite dilution of the ions $+NH_3.CH_2.COOH$ and $\cdot NH_3OH.CH_2.COO^-$ can be found as follows. Winkelblech (Z.physikal. Chem., 36, 560) gives μ_∞ for sodium glycine as 85·5. Thus for the glycine anion $NH_3OH\cdot CH_2\cdot COO^-$ μ_∞ is 85·5.- μ_{Na} = 85·5 - 50·5 = 35. Winkelblech (loc.cit.) gives for glycine hydrochloride μ_∞ = 120·4. Thus for $+NH_3.CH_2.COOH$, μ_∞ = 120·4 - μ_{Cl} = 120·4 - 75·2 = 45·2. Thus μ_∞ for glycine (anion and cation) = 35 + 45·2 = 80·2. Similarly the value of μ_∞ for β-amino propionic acid can be calculated from the results of Bork (loc. cit.) and is found to be 72·8.

Taking the molecular conductivity in each case at the dilutions considered as being ·25 and the value of μ_∞ as 75, we have the degree of dissociation $= \frac{\cdot 25}{75} = \cdot 0033$. From this we can calculate the

concentration/

concentration of the ions in an ampholyte solution as in Table 39.

Table 39.

Concentration of ampholyte.	Concentration of ions.
·20 N	2 × ·0033 × ·20 = ·00132
·10 N	2 × ·0033 × ·10 = ·00066
·05 N	2 × ·0033 × ·05 = ·00033
·025 N	2 × ·0033 × ·025 = ·00016

Thus in a ·20 N solution of glycine the concentration of ions cannot exceed ·00132 gram ions per litre. The effect of such a solution of ions on the solubility of salts can be calculated from the equation of Debye and Hückel. In its simplified form we have

$$\log \frac{S}{S_0} = \cdot 3 \left(\sqrt{\Sigma c v^2} - \sqrt{\Sigma c_0 v^2} \right)$$

The calculated values of $\log \frac{S}{S_0}$ for silver bromate in glycine solutions are shown in Table 40 and compared with the values of $\log \frac{S}{S_0}$ found by experiment

Table 40.

Concentration of Glycine	Concentration of Glycine ions	Conc. of $AgBrO_3$	$\sqrt{\Sigma c v^2}$	$\log \frac{S}{S_0}$ (calc)	$\log \frac{S}{S_0}$ (found)
·000	·00000	·00827	·1286	·0000	·0000
·025	·00016	·00827	·1292	·0002	·0124
·050	·00033	·00828	·1298	·0004	·0220
·100	·00066	·00829	·1311	·0008	·0411
·200	·00132	·00832	·1336	·0015	·0702

The/

The figures given in Column 3 of Table 40 are the values of the solubilities of $AgBrO_3$ obtained by extrapolating to the required ionic concentration the solubility curve of $AgBrO_3$ in solutions of ammonium nitrate. Almost identical values of $\log \frac{S}{S_0}$ are obtained by assuming in the calculation that the only change inthe concentration of ions is that due to the added ions of glycine. The values of $\log \frac{S}{S_0}$ by this calculation are seen to be about 1/50th of those found by experiment. Actually the probable value of $\log \frac{S}{S_0}$ to be expected on the basis that the increase is due to the moving ions of glycine is much smaller than that here estimated, and that for two reasons. Firstly, we have assumed that the conductivity of the glycine is due solely to the organic anion and cation. If any considerable portion of the current is borne by H^+ or OH^- ions, as is certainly the case, the number of ions in the solution must be smaller than we have calculated. Secondly, we have neglected the effect of the un/ionised molecules of the glycine on the solubility of the salt. These will almost certainly decrease the solubility, as non-electrolytes in general do. As these molecules are greatly more numerous than the moving ions the result to be expected is probably a decrease rather than increase in solubility. This conclusion is confirmed by the work of Hill (J.Amer. Chem.Soc., 39, 218) who found that the solubility of $AgBrO_3$ is depressed in the presence of acetic acid.

In/

In Table 41 are shown in columns 1 and 2 the normality of acetic acid and the solubility (normality) of silver bromate as found by Hill. In column 3 are given the values of the concentration of the ions of acetic acid obtained by interpolation from Kohlrausch's data.

Table 41.

Normality of Acetic acid.	Normality of Silver Bromate.	Concentration of ions of acetic acid.
.00	.00827	.0000
.05	.00824	.0019
.10	.00822	.0027
.20	.00816	.0036

It will be seen that although the concentration of added ions is 3 - 6 times as great as in the case of glycine there is a marked decrease in the solubility of silver bromate. It may thus be concluded

1. That, on the basis of the old theory it is reasonable to expect that the solubility of salts should be depressed rather than increased in the presence of ampholytes.
2. Any attempt to explain the increase in solubility as due to the moving ions of the ampholyte leads to values at least 50 times too small.

What explanation, then, can be given of the increase? It seems improbable that it is due to some chemical action such as compound formation, as the effect is of similar magnitude in such divergent cases. It/

It may be explained on Bjerrum's theory of the hybrid ion, which lowers the activity of the salt in the same way as ions in general do. As will be seen from the values of r given in Tables 3 to 37 the effect of such ions is less than half that of two univalent ions. I have not found it possible to calculate the effect which such an ion should have.

The calculation of Debye and Hückel (physikal. Z., 24, 185) proceeds from the basis that any element of volume in the vicinity of a positive ion will tend to contain more negative than positive ions, and they succeeded in calculating the average distribution of ions around a given ion and hence the potential at any point due to the ion itself and to the unequal distribution of surrounding ions which it produces. In the same way it should be possible to calculate the distribution of hybrid ions around a given ion, and hence the potential at any point due to the ion itself and to the unequal distribution of the charges which it produces. The actual effect in the case of hybrid ions will not of course be a preponderance of, say negative ions in the neighbourhood of a positive ion but the orientation of the hybrid ions so that their negative ends are towards the positive ion more frequently than the positive ends. Regarding the case from the other point of view, namely, the distribution of ions around a hybrid ion. We will have in the vicinity of the +ve charge an effect similar to that of a simple positively charged ion, but modified in two ways, Firstly, the ions are not free

to/

to take up their positions all around the +ve charge owing to the space occupied by the molecule on one side; secondly the potential at any point in the vicinity of the +ve charge is modified by the presence of the -ve charge of the hybrid ion at a fixed distance away. From either point of view, however, the result will be an unequal distribution of charges quite similar in kind to that calculated by Debye and Huckel and its magnitude may possibly be theoretically estimated at some future time.

As it seems impossible at this stage to calculate the effect of a hybrid ion on solubility of salts it is not possible to use the solubilities as a means of estimating accurately the amount of hybrid ions in solutions of ampholytes. If, however, we assume that the solubility increases are due to hybrid ions we can calculate the minimum quantities of hybrid ions by dividing the apparent"ionic strength" of the solution as determined from the solubility curves of the salt, by the ionic strength calculated for a salt solution of the equivalent strength. This has been done in Tables 42 to 68 . Another way of regarding the results is as a measure of the external field of the hybrid ions. Assuming the practically complete ionisation of the ampholyte into hybrid ions, as Bjerrum does, the ratio of the effect measured to that calculated/

calculated for a salt, will represent the ratio of the external field of the charges of the hybrid ion, to the field which is intramolecular.

The ratios r given in Tables 3 to 42 are not directly a measure of the relative effect of the hybrid ion. In order to estimate the apparent concentration of ions in the ampholyte solutions use was made of the "ionic strength" principle of G.N.Lewis and the theory of Debye & Huckel. Debye & Huckel's theory gives for the solubility of uni-univalent salt in a solution of ions

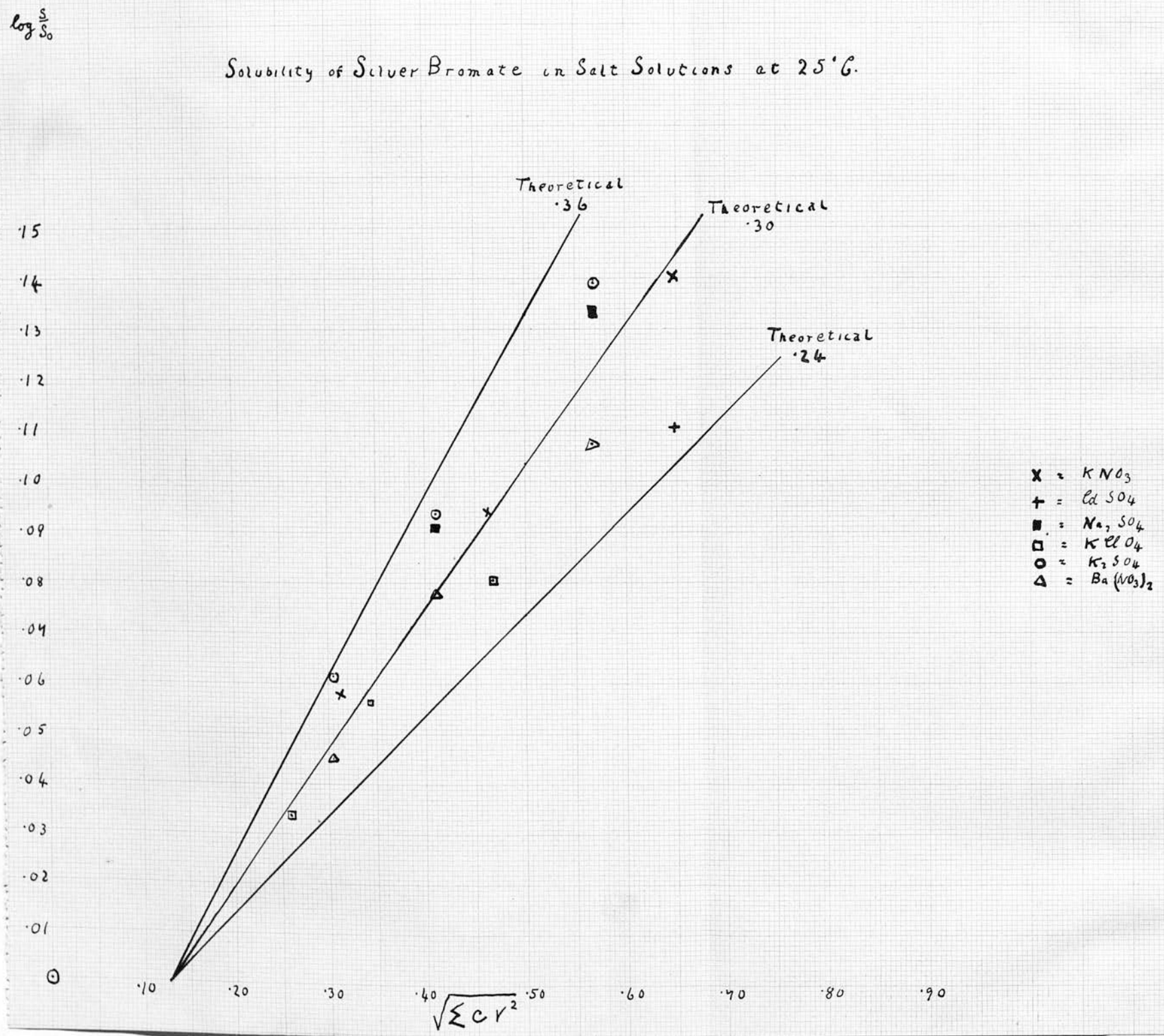
$$\log \frac{S}{S_0} = \cdot 357 \left(\sqrt{\Sigma c v^2} - \sqrt{\Sigma c_0 v^2} \right)$$

where S_0 is the molal solubility of the salt in pure water, S is the molal solubility of thesalt in a salt solution, C is the molal concentration of each ion in the solution and V is the valency of each ion. This contains the principle of Lewis, who gave the purely empirical rule that the solubility of a salt was the same in all solutions of the same"ionic strength" where the ionic strength was defined as being equal to $\frac{1}{2} \Sigma C V^2$. The factor ·357 which is the theoretical result of the calculation of Debye & Hückel is found to represent an upper limit and in practice the value is found to vary between ·36 and ·24 and indeed to fall far below this value even in reasonably dilute solutions (·10N) especially in the case of salts of the polyvalent types.

If we plot the values of $\log \frac{S}{S_0}$ against $\sqrt{\Sigma c v^2}$ we should in the ideal case obtain a straight line. In fact/

Graph I

Silver Bromate


log S/S₀
Solubility of Silver Bromate in Salt Solutions at 25°C.
Theoretical ·36
Theoretical ·30
Theoretical ·24
·15
·14
·13
·12
·11
·10
·09
·08
·04
·06
·05
·04
·03
·02
·01
0
·10
·20
·30
·40
·50
·60
·70
·80
·90
√Σcv²
X = KNO₃
+ = Cd SO₄
■ = Na₂ SO₄
□ = K Cl O₄
○ = K₂ SO₄
△ = Ba (NO₃)₂

fact using any given salt in solutions of other salts we obtain a sheaf of approximately straight lines lying chiefly within the zone bounded by the lines corresponding to the theoretical equation with the factors ·36 and ·24. The case of Silver Bromate is shown in Graph I. The solubility measurements were done by Dalton, Pomeroy and Weymouth (J.Amer.Chem.Soc., 46,63). From the graph we can find the apparent ionic strength of a given ampholyte solution by reading off the value of $\sqrt{\Sigma cv^2}$ which corresponds to the value found for $\log \frac{S}{S_0}$. Hence we find the value of Σcv^2 for the solution and by subtracting from this the value of Σcv^2 for the ions of silver bromate ($\Sigma c_s v^2$) we get the value of Σcv^2 for the ampholyte ($\Sigma c_a v^2$). In this way we find the values shown in Table 42 obtained from the central line (Theoretical = ·30)

Table 42.

Silver Bromate in Glycine.

a	$\frac{S}{S_0}$	$\log \frac{S}{S_0}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$c_s v^2$	$c_a v^2$	γ
·000	1·000	·000	·128	·0165	·0165	·0000	–
·025	1·033	·014	·181	·0328	·0171	·0157	·31
·050	1·052	·022	·209	·0437	·0174	·0263	·26
·100	1·099	·041	·278	·0773	·0182	·0591	·29
·200	1·175	·070	·383	·1470	·0194	·1276	·32
						Mean	·30

In the last column headed γ is given the value of the ionic strength thus found divided by the ionic strength calculated for a uni-univalentsalt of equivalent concentration. This means that the apparent concentration/

concentration of ions in the glycine solutions is about ·30 of that which it would have on the assumption that the hybrid ion has an effect equal to that of two univalent ions of opposite charge completely separated from each other.

In view of the anomalous effect in the case of the silver and lead salts mentioned before (p.43) I have not carried out this calculation in the case of the other amino carboxylic acids with silver bromate and lead bromide. The results of the calculation in the cases of the other salts are given in Tables 43 to 68 .

Table 43.

Silver Bromate in Taurine.

a	$\frac{S}{S_o}$	$\log\frac{S}{S_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·128	·0165	·0165	·0000	–
·025	1·018	·0076	·155	·0240	·0168	·0072	·14
·050	1·028	·0119	·171	·0292	·0170	·0122	·12
·100	1·046	·0195	·200	·0400	·0173	·0227	·11
·200	1·083	·0346	·254	·0645	·0179	·0466	·12
						Mean =	·12

Table 44.

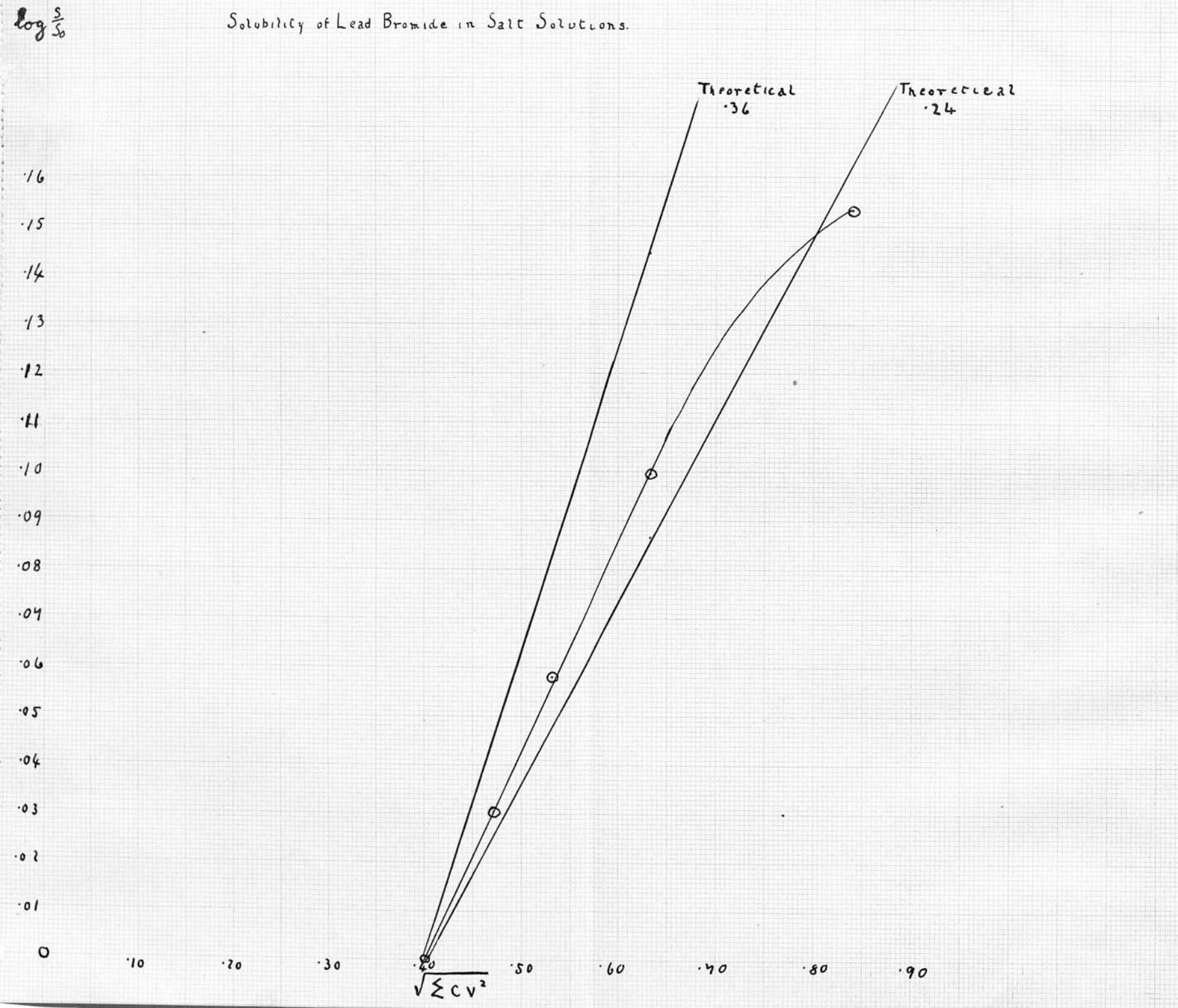
Silver Bromate in Taurobetaine.

a	$\frac{S}{S_o}$	$\log\frac{S}{S_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·128	·0165	·0165	·0000	–
·025	1·018	·0076	·155	·0240	·0168	·0072	·14
·050	1·031	·0132	·176	·0310	·0171	·0139	·14
·100	1·053	·0224	·210	·0441	·0174	·0267	·13
·200	1·094	·0391	·270	·0729	·0181	·0548	·14
						Mean =	·14

In the case of lead bromide the only results available/

Graph 2

Lead Bromide


log S/S₀
Solubility of Lead Bromide in Salt Solutions.
Theoretical
·36
Theoretical
·24
·16
·15
·14
·13
·12
·11
·10
·09
·08
·07
·06
·05
·04
·03
·02
·01
0
·10
·20
·30
·40
·50
·60
·70
·80
·90
√Σ C V²

available for plotting $\log \frac{S}{S_0}$ against $\sqrt{\Sigma cv^2}$ are these which I obtained for the solubility of lead bromide in ammonium nitrate. The calculations are shown in Table 45 and the results are plotted in Graph 2 . This graph was used in finding the values of $\sqrt{\Sigma cv^2}$ given in Tables 46 and 47 .

Table 45 .

Lead Bromide in Ammonium Nitrate.

a	S	$\frac{S}{S_0}$	$\log \frac{S}{S_0}$	Σcv^2	$\sqrt{\Sigma cv^2}$
·000	·0267	1·000	·0000	·1602	·400
·025	·0287	1·073	·0306	·2224	·471
·050	·0303	1·143	·0580	·2818	·531
·100	·0335	1·258	·0997	·4010	·633
·200	·0380	1·423	·1532	·6380	·829

Theoretical for ·10 M:- $\frac{1}{2} \log \frac{S}{S_0} = ·36(·633-·400) = ·0722$

$\log \frac{S}{S_0} = ·1444$

or $\frac{1}{2} \log \frac{S}{S_0} = ·24(·633-·400) = ·0432$

$\log \frac{S}{S_0} = ·0864$

Table 46.

Lead Bromide in Taurine.

a	$\frac{S}{S_0}$	$\log \frac{S}{S_0}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·400	·160	·160	·000	-
·025	1·034	·0145	·434	·188	·166	·022	·44
·050	1·060	·0253	·460	·212	·170	·042	·42
·100	1·105	·0433	·501	·251	·177	·084	·42
·200	1·187	·0745	·574	·330	·190	·140	·35
						Mean =	·41

Table 47.

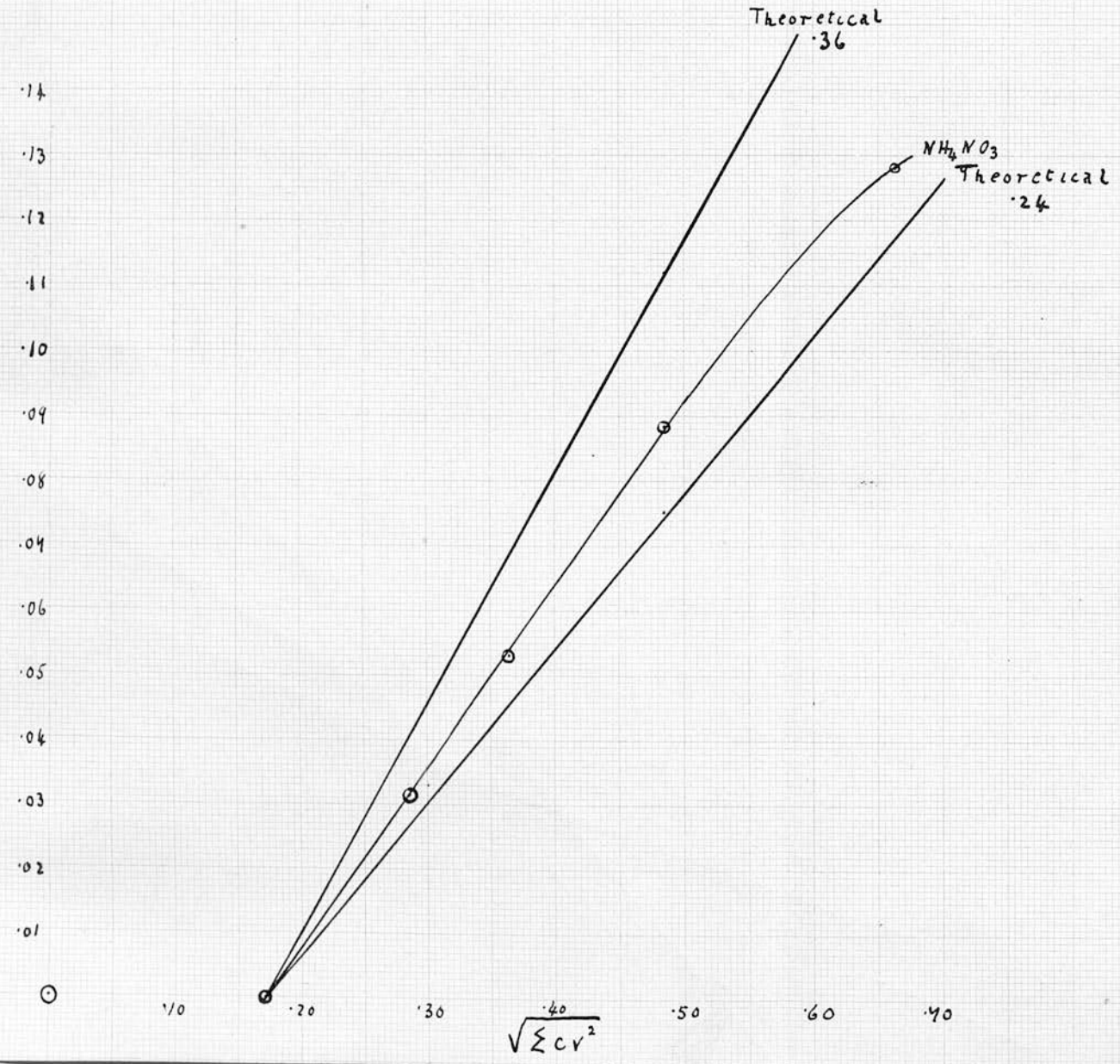
Lead Bromide in Taurobetaine.

a	$\frac{S}{S_0}$	$\log \frac{S}{S_0}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·400	·160	·160	·000	-
·025	1·034	·0145	·434	·188	·166	·022	·44
·050	1·067	·0282	·466	·217	·171	·046	·46
·100	1·134	·0546	·528	·279	·180	·099	·45
·200	1·224	·0878	·605	·366	·196	·176	·44
						Mean =	·45

Graph 3

Thallous Thiocyanate

Solubility of Thallous Thiocyanate in Salt Solutions at 25° C.



The only figures available for plotting $\log \frac{S}{S_0}$ against $\sqrt{\Sigma cv^2}$ in the case of thallous thiocyanate are those obtained bymyself for the solubility of this salt in ammonium nitrate. The calculations are shown in Table 48 and the results are plotted in Graph 3 . This graph was used in finding the values of $\sqrt{\Sigma cv^2}$ given in Tables 49 to 54 .

Table 48.

Thallous Thiocyanate in Ammonium Nitrate.

a	s	$\frac{s}{s_0}$	$\log \frac{s}{s_0}$	Σcv^2	$\sqrt{\Sigma cv^2}$
·000	·0147	1·000	·0000	·0294	·171
·025	·0158	1·075	·0315	·0816	·286
·050	·0166	1·130	·0531	·1332	·365
·100	·0180	1·1225	·0881	·2360	·485
·200	·0202	1·374	·1380	·4404	·664

Theoretical for ·10 N:-

$\log \frac{S}{S_0}$ = ·36 (·485 - ·171) = ·1120

or $\log \frac{S}{S_0}$ = ·24 (·485 - ·171) = ·0754.

Table 49.

Thallous Thiocyanate in Glycine.

a	$\frac{s}{s_0}$	$\log \frac{s}{s_0}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·171	·0294	·0294	·0000	-
·025	1·014	·0060	·194	·0376	·0298	·0078	·16
·050	1·027	·0115	·213	·0454	·0302	·0152	·15
·100	1·041	·0174	·234	·0547	·0306	·0241	·12
·200	1·075	·0215	·250	·0625	·0316	·0309	·08
						Mean =	·13

Table 50.

Thallous Thiocyanate in βamino propionic acid

a	$\frac{s}{s_0}$	$\log \frac{s}{s_0}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·171	·0294	·0294	·0000	-
·025	1·014	·0060	·194	·0376	·0298	·0078	·16
·050	1·034	·0145	·244	·0595	·0304	·0291	·29
·100	1·062	·0271	·270	·0729	·0312	·0417	·21
·200	1·113	·0464	·337	·1136	·0328	·0808	·20
						Mean =	·21

Table 51.

Thallous Thiocyanate in γ amino butyric acid.

a	$\frac{S}{S_o}$	$\log \frac{S}{S_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·171	·0294	·0294	·0000	–
·025	1·024	·0115	·213	·0454	·0302	·0152	·30
·050	1·048	·0203	·245	·0600	·0308	·0292	·29
·100	1·102	·0422	·343	·1176	·0324	·0852	·42
·200	1·184	·0734	·434	·1880	·0348	·1532	·38
						Mean =	·48

Table 52.

Thallous Thiocyanate in δ amino valeric acid.

a	$\frac{S}{S_o}$	$\log \frac{S}{S_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·00	1·000	·0000	·171	·0294	·0294	·0000	–
·10	1·095	·0395	·314	·0986	·0322	·0664	·33
·20	1·156	·0630	·396	·1568	·0340	·1228	·31
						Mean =	·32

Table 53.

Thallous Thiocyanate in Taurine.

a	$\frac{S}{S_o}$	$\log \frac{S}{S_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·171	·0294	·0294	·0000	–
·025	1·014	·0060	·194	·0376	·0298	·0078	·16
·050	1·034	·0145	·244	·0595	·0304	·0291	·29
·100	1·068	·0286	·274	·0751	·0314	·0437	·22
·200	1·130	·0531	·362	·1310	·0332	·0978	·24
						Mean =	·23

Table 54.

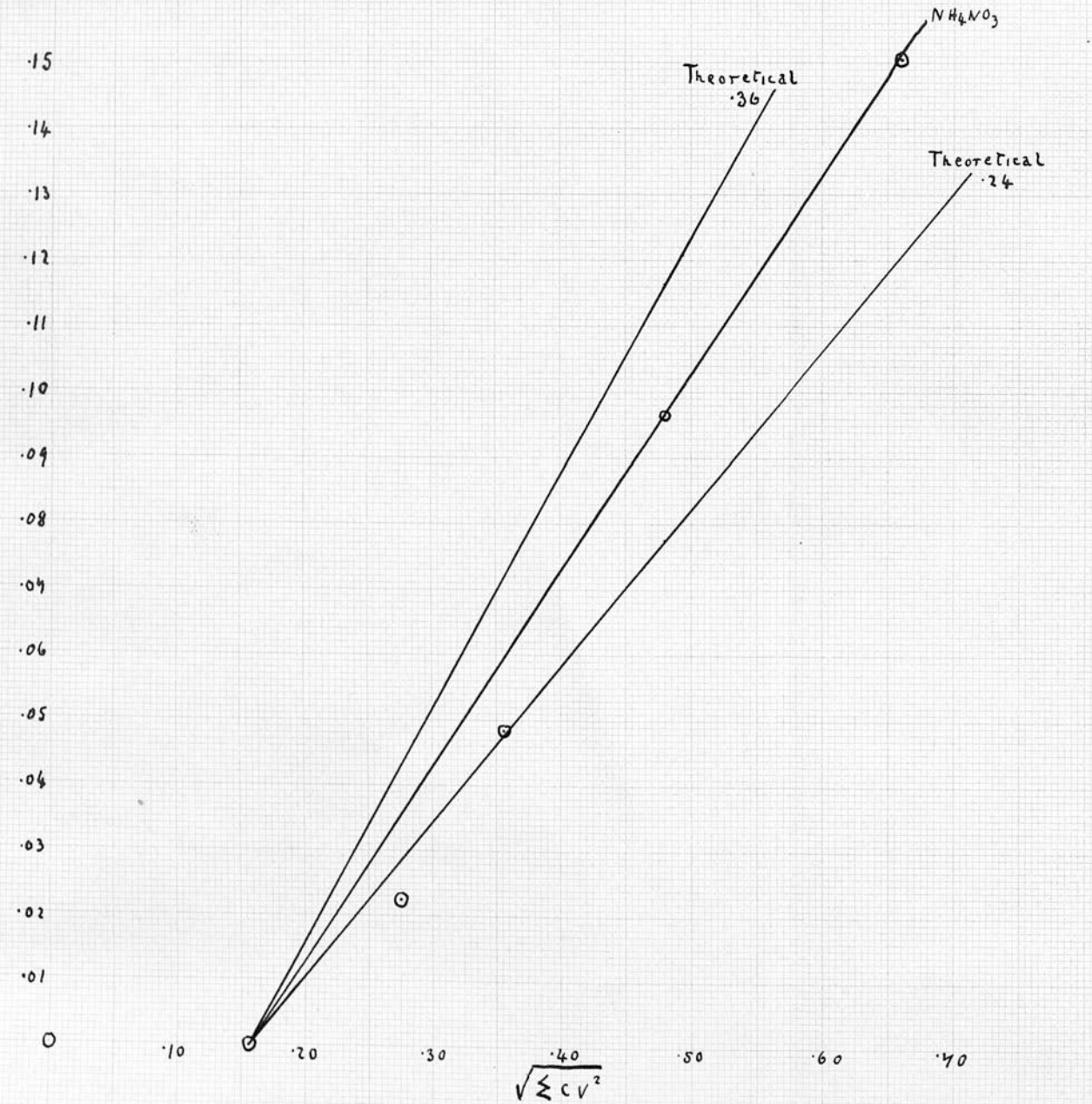
Thallous Thiocyanate in Taurobetaine.

a	$\frac{S}{S_o}$	$\log \frac{S}{S_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·171	·0294	·0294	·0000	–
·025	1·020	·0086	·203	·0412	·0300	·0112	·22
·050	1·041	·0174	·234	·0548	·0306	·0242	·24
·100	1·088	·0367	·303	·0918	·0320	·0598	·30
·200	1·170	·0682	·416	·1731	·0344	·1387	·34
						Mean =	·27

Graph 4

Thallous Bromate

Solubility of Thallous Bromate in Salt Solutions



The only figures available for plotting $\log \frac{S}{S_o}$ against $\sqrt{\Sigma cv^2}$ in the case of Thallous Bromate are these which I have obtained for the solubility of this salt in Ammonium Nitrate. The calculations are shown in Table 55 and the results are plotted in Graph 4 . This graph was used in finding the values of $\sqrt{\Sigma cv^2}$ given in Tables 56 to 61.

Table 55.

Solubility of Thallous Bromate in Ammonium Nitrate.

a	S	$\frac{S}{S_o}$	$\log \frac{S}{S_o}$	Σcv^2	$\sqrt{\Sigma cv^2}$
·000	·0124	1·000	·0000	·0248	·157
·025	·0130	1·053	·0224	·0760	·276
·050	·0138	1·118	·0483	·1274	·357
·100	·0155	1·249	·0965	·2310	·480
·200	·0176	1·414	·1504	·4352	·660

Theoretical for ·10 N:-

$$\log \frac{S}{S_o} = \cdot 36\ (\cdot 480 - \cdot 157) = \cdot 1163$$

or

$$\log \frac{S}{S_o} = \cdot 24\ (\cdot 480 - \cdot 157) = \cdot 0775$$

Table 56.

Thallous Bromate in Glycine.

a	$\frac{S}{S_o}$	$\log \frac{S}{S_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·157	·0247	·0247	·0000	–
·025	1·014	·0060	·175	·0306	·0260	·0046	·09
·050	1·023	·0098	·190	·0351	·0274	·0087	·09
·100	1·059	·0249	·238	·0567	·0310	·0257	·13
·200	1·114	·0468	·313	·0980	·0352	·0628	·16
						Mean =	·12

Table 57.

Thallous Bromate in β amino propionic acid.

a	$\frac{S}{S_o}$	$\log \frac{S}{S_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·157	·0247	·0247	·0000	–
·025	1·014	·0060	·175	·0306	·0251	·0055	·11
·050	1·024	·0103	·192	·0369	·0254	·0115	·13
·100	1·070	·0294	·256	·0655	·0265	·0390	·19
·200	1·134	·0546	·340	·1183	·0281	·0898	·22
						Mean =	·16

Table 58.

Thallous Bromate in γ amino butyric acid.

a	$\frac{s}{s_o}$	$\log \frac{s}{s_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·157	·0247	·0247	·0000	–
·025	1·016	·0068	·179	·0320	·0252	·0068	·14
·050	1·028	·0119	·196	·0384	·0255	·0129	·13
·100	1·076	·0319	·264	·0697	·0266	·0433	·21
·200	1·150	·0607	·360	·1296	·0285	·1009	·25
						Mean =	·18

Table 59.

Thallous Bromate in δ amino valeric acid.

a	$\frac{s}{s_o}$	$\log \frac{s}{s_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·00	1·000	·0000	·157	·0247	·0247	·0000	–
·10	1·066	·0278	·250	·0625	·0264	·0361	·18
·20	1·133	·0542	·338	·1142	·0281	·0861	·21
						Mean =	·20

Table 60.

Thallous Bromate in Taurine.

a	$\frac{s}{s_o}$	$\log \frac{s}{s_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·157	·0247	·0247	·0000	–
·025	1·007	·0029	·167	·0279	·0249	·0030	(·06)
·050	1·014	·0060	·175	·0306	·0251	·0055	(·06)
·100	1·051	·0216	·228	·0520	·0260	·0260	·13
·200	1·096	·0399	·290	·0841	·0271	·0570	·14
						Mean =	·13

Table 61.

Thallous Bromate in Taurobetaine.

a	$\frac{s}{s_o}$	$\log \frac{s}{s_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·157	·0247	·0247	·0000	–
·025	1·009	·0037	·169	·0286	·0250	·0036	(·07)
·050	1·027	·0115	·195	·0380	·0254	·0126	·13
·100	1·059	·0249	·240	·0576	·0262	·0314	·15
·200	1·113	·0464	·312	·0974	·0276	·0698	·17
						Mean =	·15

Graph 5

Calcium Iodate

Again in the case of Calcium Iodate the only available values for plotting $\log \frac{S}{S_0}$ against $\sqrt{\Sigma cv^2}$ are those I obtained for the solubility of this salt in ammonium nitrate solutions. The calculations are shown in Table 62 and the results are plotted in Graph 5 . This graph was used in finding the values of $\sqrt{\Sigma cv^2}$ given in Tables 63 to 68 .

Table 62.

Calcium Iodate in Ammonium Nitrate.

a	S	$\frac{S}{S_0}$	$\log \frac{S}{S_0}$	Σcv^2	$\sqrt{\Sigma cv^2}$
·000	·0079	1·000	·0000	·0474	·218
·025	·0088	1·108	·0444	·1028	·321
·050	·0095	1·196	·0778	·1570	·396
·100	·0107	1·348	·1297	·2642	·514
·200	·0125	1·580	·1987	·4750	·689

Theoretical for ·10 M :-

$$\tfrac{1}{2} \log \frac{S}{S_0} = \cdot 36 \, (\cdot 514 - \cdot 218) = \cdot 1066$$

$$\therefore \log \frac{S}{S_0} = \cdot 2132$$

or

$$\tfrac{1}{2} \log \frac{S}{S_0} = \cdot 24 \, (\cdot 514 - \cdot 218) = \cdot 0710$$

$$\therefore \log \frac{S}{S_0} = \cdot 1420$$

Table 63.

Calcium Iodate in Glycine.

a	$\frac{S}{S_0}$	$\log \frac{S}{S_0}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·218	·0474	·0474	·0000	–
·025	1·019	·0080	·237	·0562	·0483	·0079	·16
·050	1·045	·0191	·261	·0681	·0495	·0186	·18
·100	1·091	·0378	·305	·0930	·0517	·0413	·20
·200	1·205	·0809	·404	·1632	·0571	·1061	·26
						Mean =	·20

Table 64.

Calcium Iodate in β amino propionic acid.

a	$\frac{S}{S_o}$	$\log \frac{S}{S_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·218	·0474	·0474	·0000	–
·025	1·027	·0115	·245	·0600	·0486	·0114	·23
·050	1·051	·0216	·268	·0718	·0498	·0220	·22
·100	1·096	·0399	·310	·0961	·0519	·0442	·22
·200	1·203	·0803	·403	·1624	·0570	·1054	·26
						Mean =	·23

Table 65.

Calcium Iodate in γ amino butyric acid.

a	$\frac{S}{S_o}$	$\log \frac{S}{S_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·218	·0474	·0474	·0000	–
·025	1·027	·0115	·245	·0600	·0486	·0114	·23
·050	1·053	·0224	·270	·0729	·0500	·0229	·23
·100	1·089	·0371	·304	·0924	·0516	·0408	·20
·200	1·209	·0823	·407	·1656	·0573	·1083	·27
						Mean =	·23

Table 66.

Calcium Iodate in δ amino valeric acid.

a	$\frac{S}{S_o}$	$\log \frac{S}{S_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·00	1·000	·0000	·218	·0474	·0474	·0000	–
·10	1·118	·0483	·330	·1089	·0530	·0559	·28
·20	1·220	·0864	·416	·1731	·0578	·1153	·29
						Mean=	·28

Table 67.

Calcium Iodate in Taurine.

a	$\frac{S}{S_o}$	$\log \frac{S}{S_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·218	·0474	·0474	·0000	–
·025	1·013	·0055	·231	·0534	·0480	·0054	(·11)
·050	1·038	·0161	·255	·0656	·0492	·0158	(·16)
·100	1·089	·0371	·305	·0930	·0516	·0414	·21
·200	1·165	·0664	·372	·1383	·0552	·0831	·21
						Mean =	·17

Table 68.

Calcium Iodate in Taurobetaine.

a	$\frac{S}{S_o}$	$\log \frac{S}{S_o}$	$\sqrt{\Sigma cv^2}$	Σcv^2	$\Sigma c_s v^2$	$\Sigma c_a v^2$	Y
·000	1·000	·0000	·218	·0474	·0474	·0000	–
·025	1·007	·0029	·225	·0506	·0477	·0029	·06
·050	1·013	·0055	·232	·0539	·0480	·0059	·06
·100	1·038	·0161	·256	·0656	·0492	·0164	·08
·200	1·076	·0317	·291	·0847	·0510	·0337	·08
						Mean =	·07

The results obtained in Tables 42 to 68 have been summarised in Table 69 which gives the mean values of γ found for each ampholyte with each salt.

Summary Table 69.

1	2	3	4	5	6
Ampholyte	$AgBrO_3$	$PbBr_2$	TlSCN	$TlBrO_3$	$Ca(IO_3)_2$
Glycine	·30	–	·13	·12	·20
amino propionic acid	–	–	·21	·16	·23
amino butyric acid	–	–	·35	·18	·23
amino valeric acid	–	–	·32	·20	·28
Taurine	·12	(·41)	·23	·13	·17
Taurobetaine	·14	(·45)	·27	·15	(·07)

From an examination of Table 69 no great regularities appear. In Columns 4, 5 and 6 there appears to be a regular increase in the effect with increasing length of chain which is most marked in the case of TlSCN. Omitting the three figures shown in brackets the figures are compatible with the theory that they express an effect due to a common physical cause. On the other hand, the agreement between the figures is not sufficiently close to make it certain that this is so; there remains the possibility that the effect is due to some specific chemical action in each case, the extreme cases of which (with lead bromide) have been omitted from the table. It must be remembered that the method of analyses and the quantities/

quantities of ampholyte available rendered high accuracy impossible and small errors in the analyses will result in considerable errors in the result especially at the lower concentrations. The errors are likely to be at their maximum in the cases of γ and δ acids which were in very small quantities. There remains a series of numbers varying from ·12 to ·27 expressing the relative effect of the hybrid ions of the α and β types on the solubility of salts.

The assumption that this effect is due to the electric charges on the hybrid ion and not to the chemical forces may be made, and its consequences found. The first consequence of this assumption is that the solubility of ampholytes should be increased in the presence of salts. This is in agreement with the observation of Bjerrum on the solubility of Methyl Orange in salt solutions (p. 6). An examination of the literature revealed the existence of some other data not mentioned by Bjerrum. This is the work of Pfeiffer (Ber., 48, 1938) on the solubility of glycine in solutions of the chlorides and bromides of the alkalis and alkaline earths. The case of strontium chloride is the only one in which Pfeiffer gives the effect of different concentrations of the salt. In Table 70 I have calculated the results of Pfeiffer in terms of the new theory. In columns 1 and 2 are the concentrations of salt used and ampholyte saturating/

Graph 6

Glycine in Salt Solutions

saturating the solution expressed as moles per litre; in column 3 is given Σcv^2 for salt + ampholyte assuming that glycine contributes 1/5th of the theoretical ionic strength for a salt. In column 4 is given $\sqrt{\Sigma cv^2}$ and in column 5, $\log \frac{S}{S_0}$ calculated from the values given in column 2.

Table 70.

Solubility of Glycine in Strontium Chloride.

a	s	Σcv^2	$\sqrt{\Sigma cv^2}$	$\log \frac{S}{S_0}$
·00	2·62	1·05	1·02	·000
·25	2·84	2·63	1·62	·036
·50	3·11	4·24	2·06	·075
1·00	3·47	6·39	2·53	·123
2·00	4·38	13·75	3·71	·227

In graph 6 $\log \frac{S}{S_0}$ is plotted against $\sqrt{\Sigma cv^2}$. The circles represent the plots for strontium chloride while the squares represent the plots for other salts. It will be seen that the curve of solubility of glycine in strontium chloride does not depart from a straight line any more than may be expected in such concentrated solutions. The simplest way in this case to calculate the magnitude of the ratio of the observed effect for glycine to that calculated for a uni-univalent salt is to divide the slope of the curve by the theoretical slope of ·36 (the factor α in Debye and Hückels equation). From the line AB representing the average slope we have $\alpha = \frac{·10}{1·08} = ·0926$. Hence γ for glycine is ·0926 - ·36 = ·25. Thus the solubility increase is ·25 of that which would occur if glycine were a salt without internal compensation of/

of the charges. This result, taking into account the high concentration of the solutions is in excellent agreement with the theory of the hybrid ions; and it is unnecessary to introduce any hypothesis of compound formation in solution as Pfeiffer has done. The isolation of crystalline compounds such as $SrCl_2 \cdot 2NH_2 \cdot CH_2 \cdot COOH \cdot 3H_2O$ is possibly to be explained as due to "ampholyte of crystallisation". This is more probable as the ampholyte is now assumed to be a very highly polar substance like water, and moreover the compounds occur with varying proportions of ampholyte and water recalling the various hydrates of the salts.

Another consequence of the conclusion that the hybrid ion lowers the activity of ions in solution is that it must lower its own activity. i.e. in a solution of an ampholyte there should be a departure from the laws of solution of the same type as the osmotic deviations of electrolytes. Now a hybrid ion lowers the activity of a given simple ion by an amount equal to .20 - .40 of the amount by which an ordinary ion lowers the activity of that given ion. Similarly a given ordinary ion will lower the activity of the hybrid ion by .20 - .40 of the amount by which it lowers the activity of another ordinary ion. Thus we may expect that the hybrid ion will lower the activity of another hybrid ion by an amount equal to $.20^2$ to $.40^2$ of that by which an ordinary ion lowers the activity of other ordinary ions. Hence in a solution of an ampholyte the osmotic properties should/

Graph 7

F. P. of Glycine Solutions

should assume values less than the theoretical i.e. there should be an apparent association increasing with the concentration (just as the apparent dissociation of a salt diminishes with concentration). The magnitude of this effect should be between $\cdot 20^2$ and $\cdot 40^2$ of that of an ordinary ion i.e. between $\cdot 04$ and $\cdot 16$ of the effect of an ordinary ion. An accurate calculation of the magnitude of the osmotic deviation is not possible from the figures I have obtained for the activities but an approximate verification of this conclusion can be obtained. In graph 7 I have plotted against the concentration, the difference between the F.P. depressions found and those calculated from the formula $\Delta = 1 \cdot 86 \times C$ where C is the total concentration of the species present. The figures for NaCl are typical for all uni-univalent salts. They are taken from Jahn (Z.physikal.Chem., 50, 136); the figures for glycine are from Roth (Z.physikal.Chem., 43, 558) It will be seen that the values for glycine are about 1/10 of those for a uni-univalent salt. This is in agreement with the argument outlined above as will be seen from the calculation below.

For $\cdot 10$ N salt we have the equation of Debye & Hückel

$$1 - \phi = \cdot 27 \left(\sqrt{\Sigma c v^2} \right)$$

$$= \cdot 27 \left(\sqrt{\cdot 20} \right)$$

$$= \cdot 120$$

Now in the case of $\cdot 10$N glycine, assuming that $\Sigma c_a v^2$

for/

for ·10N Glycine = ·03

$$1 - \varphi = \cdot 27\sqrt{\cdot 03}$$

$$= \cdot 047$$

·047 represents the deviation in a solution of ordinary ions of strength $\Sigma cv^2 = \cdot 03$ but the fact that we are dealing with a hybrid ion in such a solution will reduce the deviation to ·20 - ·40 of this value, - say to ·01

$$\text{i.e.} \quad 1 - \varphi = \cdot 01$$

and the ratio of the deviations $\frac{1 - \varphi(\text{ampholyte})}{1 - \varphi(\text{salt})} = \frac{1}{12.}$

or approximately the value shown by Graph 7 .

The occurrence of these effects (the increase in solubility of ampholytes in salts and the apparent association of ampholytes in solution) and the approximate calculation of their magnitude are a striking confirmation of the hypothesis that the solubility increases I have measured are really the effect of the hybrid ion.

GENERAL CONCLUSION.

The chemical properties of the amino acids strongly support the hybrid ion formula. Erlenmeyer and Sigel (Annalen, 176, 350) first pointed out that the properties of a carboxylic acid appeared in glycine only when a strong acid was added and they explained this/

this by the cyclic structure. Ostwald (J.prakt.Chem., 32,369) pointed out that the conductivity behaved on dilution like that of a salt. Marckwald, Neumark and Stelzner (Ber.,24,3279) showed that unlike all other aliphatic amines, the amino acids do not react with mustard oils; but that this reaction occurs readily when alkali is added. This strongly indicates that only in the presence of alkali does the amino group appear. The evidence of the formaldehyde titration of amino acids has been mentioned (p. 5); and the strongest chemical evidence is that of the much more plausible values of K_A and K_B on the hybrid ion theory. All the evidence concerning polar series shows that NH_2 as a substituent occupies a position in the region of Cl or OCH_3 in these series for all properties except the dissociation constants of acids, in which case the amino group has an effect about 1 million times that which its neighbours in the series have . On the other hand Tilden and Foster found that glycine reacts with nitrosyl chloride as if it contained an amino group. The bulk of the evidence thus favours the cyclic or salt-like structure for amino acids. But in the light of the modern theory of salt solutions, this salt must be completely dissociated, i.e. it occurs completely as the hybrid ion. The occurrence of the anion and cation are merely evidences of hydrolysis of the salt. In this paper the effect of various ampholytes on the solubility of/

of several salts has been determined and the apparent "ionic strength" of the ampholyte solutions has been determined. A measure has thus been obtained of the external field of force of the hybrid ions of these ampholytes. It appears that the general effect is of the order of ·10 to ·20 of the total effect to be expected from two univalent charges on separate molecules. The assumption that the effect should increase with separation of the charges does not mean that it should increase with the length of the chain; for a consideration of the possible positions in space of the N and O atoms shows that, when we take the most favourable direction of valency bonds at the tetrahedral angle, the charges can approach each other more closely in the β and γ acids than is possible in the α acid. Such a configuration will no doubt be favoured by the electrical attraction between the charges. On the other hand it is probable that the sameforces which separate the ions of a salt in solution will tend to stretch out the hybrid ions until the charges are so far apart as possible, so that an increase of the solubility effect with length of chain remains a possibility. This point cannot be considered definitely settled although the figures given in Table 69 probably indicate that the effect does increase with the length of the chain.

SUMMARY.

1. Bjerrum's theory that ampholytes consist largely of hybrid ions suggests that ampholytes should lower the activity of salts in solution and that salts should lower the activity of ampholytes in solution.

2. A series of six ampholytes has been prepared and their influence on the solubility of five salts has been determined at varying concentrations.

3. In all cases an increase in solubility is found. Assuming that this is due to the effect of the hybrid ion, as ion, the "ionic strength" of the ampholyte solutions has been calculated.

4. The general magnitude of the ionic strength is .10 to .20 of that calculated for infinite separation of the charges. In some cases specific effects are added.

5. By the use of this figure the increase in solubility of glycine in presence of strontium chloride and the apparent polymerisation of glycine in aqueous solution, noticed by other authors, have been explained as an effect of the hybrid ion.